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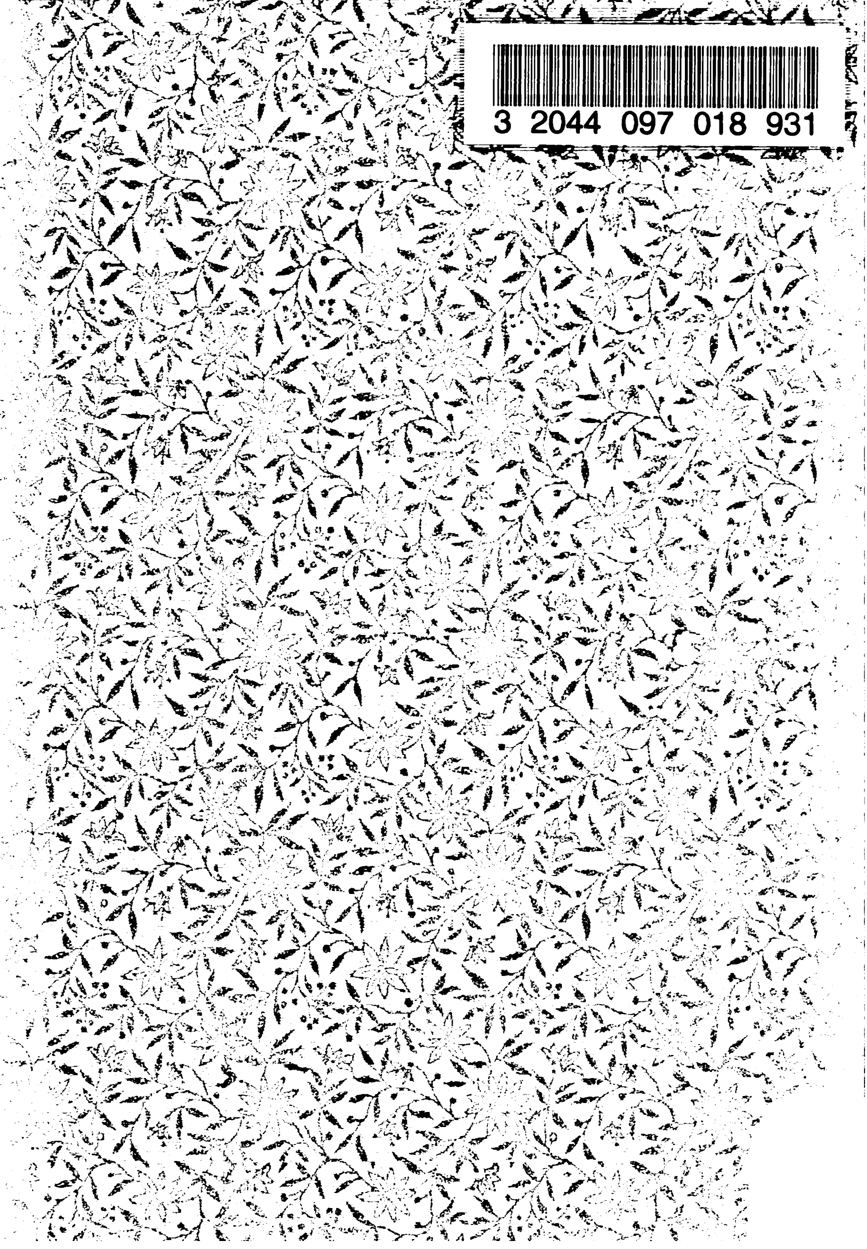

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AN

OUTLINE

OF

LECTURE NOTES

ON

GENERAL CHEMISTRY

BY

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I.

THE NON-METALS

Second Revised Edition

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PREFACE.

This little volume is not intended to take the place of the student's notes, but simply to serve as a basis of notes to be taken on a first course of experimental lectures on general chemistry. By its use the student will be relieved from the most irksome and least valuable part of his note-taking, and left free to give closer attention to the lecture and its accompanying experiments.

In this edition the text has been carefully revised and some new matter added to the appendix.

To those who have found the book useful in the past, it will, I trust, prove still more helpful in its revised form.

J. T. S.

GENERAL PRINCIPLES.

1. The Departments of Natural Science. Natural Science is a classified knowledge of Nature. Its departments may be divided with reference to their objects of study into two groups:—

a. *General Sciences,* which consider matter as such and the changes or phenomena which it undergoes. To this group belong *Chemistry* and *Physics;* they study the progress, results and causes of material phenomena and the relations which exist between them. Their field for study is as broad as the distribution of matter and the laws which they discover are *universal* laws, applicable wherever matter is found—in plant, animal or mineral; on our earth, the sun or the remotest star. b. *Special Sciences,* which are limited to certain groups or classes of objects and consider their external form or internal structure. Such are Botany, Mineralogy, Geology, Zoology.

2. The province of Chemistry distinguished from that of Physics. All natural phenomena may be divided into two classes: a. Those which consist in a loss of identity in the substance concerned. b. Those which are unaccompanied by such change of substance. The former are called *chemical* phenomena; the latter *physical.*

3. **Indestructibility of matter.** The destruction of matter which seems to occur in many chemical changes is only apparent; matter is indestructible, and, although it may be changed into an invisible gas, still has weight.

Whenever a chemical change occurs, the amount (weight) of matter after the change is the same as before.

4. **Elements and Compounds.** Most substances can be resolved by some process into two or more substances which differ essentially from the original substance and from each other; *i. e.,* most substances are *compounds.* But about seventy different kinds of matter are known, which have thus far defied all attempts at further resolution; these are, in consequence, called simple substances or *Elements.*

The elements as a class are distinguished from compounds in no way except by simplicity of composition.

All compounds are made up of these elements combined in various ways. Their composition may be determined either by *analysis* or by *synthesis.*

5. **Laws of definite and of multiple proportions.** Innumerable analyses and syntheses of compounds prove that:—a. Any given compound always contains the same elements combined in the same proportion by weight. b. When two or more compounds consist of the same elements united in differ-

ent proportions, the varying proportions of each element can be expressed by simple ratios.

6. **Effect of Heat.** a. *Expansion.* In general, all substances increase in volume when heated, and contract when cooled. Advantage is taken of the expansion of liquids (mercury or alcohol) in the construction of *thermometers.* Three thermometric scales are in use : Centigrade, Fahrenheit and Réaumer. (See p. 81.) The centigrade scale is employed in most scientific work.

Solids expand less than liquids and both much less than gases. The rate of expansion of solids and liquids is specific for each substance ; that of gases is the same for all and is quantitatively expressed by the

Law of Charles: When the pressure is constant, every gas expands $\frac{1}{273}$ part of its volume at 0° C. for every increase in temperature of 1° C. (See p. 83.)

b. *Rise of Temperature* usually attends the application of heat to a body (see however c.). The same amount of heat may produce very different changes of temperature in equal weights of different bodies. *Specific heat ; unit of heat.*

c. *Change of state of aggregation.* (1.) Passage from the solid to the liquid state, and conversely, occurs at a temperature which is in general specific and invariable for each substance—the *melting point.* During the change the temperature remains constant. *Latent heat of fusion ; freezing mixtures.*

(2.) Passage from the liquid state to that of gas or vapor takes place at all temperatures to some extent (evaporation) ; but most rapidly at the *boiling* [illegible] compounds [illegible] the formation of

[illegible] of matter [illegible] the resolution of

[illegible] as the mi[illegible] of all smaller

[illegible] atoms may be

[illegible] of molecules.

They are the smallest particles of elements which can take part in chemical changes; are for the most part incapable of existence in the free state, but are generally found in combination with other atoms, either of the same kind or of different kinds. A molecule consists of atoms held together by *chemical affinity.* In an element, the molecules consist of atoms of the same kind, or, in a few cases of one atom. In a compound the molecules are made up of atoms of different kinds.

Atoms differ from each other in weight, in chemical affinity and in valence.

a. The *atomic weight* of an element is the relative weight of its atom referred to that of the hydrogen atom as unity. *Molecular weight* is the sum of the weights of the atoms which compose the molecule.

b. The *chemical affinity* which atoms show and which determines the energy of chemical action and the stability of compounds, varies widely in intensity in different atoms.

c. *Valence* is the property of an atom by virtue of which it can hold in combination a certain *number* of other atoms. An atom exhibiting the least power in this respect is called a *monad.* An atom capable of holding in combination two monads or their equivalent is called a *dyad;* three, a *triad;* four, a *tetrad;* five, a *pentad;* &c.; or, the atoms are said to be *univalent, bivalent, trivalent, quadrivalent, quinquivalent,* &c., respectively.

The chemical properties of matter depend entirely upon the atoms of which it is composed and their relations to each other.

11. Chemical Notation. *An atom* is represented by a *symbol,* consisting of the first letter of its English or Latin name, or, when this would produce confusion, of the first with another letter of the name added. A number of atoms of the same kind is indicated by placing an Arabic numeral before the symbol, or an exponent to the right. Valence is denoted by an exponent of Roman numerals or by dashes.

Molecules are represented by grouping together the symbols of their constituent atoms. Such a group of symbols is called the *formula* of the molecule or substance. When it indicates simply the kind and number of atoms (qualitative and quantitative composition), it is an *empirical formula;* when it shows the mutual relations of the atoms, it is a *graphic* or *constitutional formula.* Molecular formulas are multiplied by coefficients or by Arabic exponents placed outside an enclosing parenthesis.

12. Chemical Nomenclature. That of the *elements* follows no special rule. The names often indicate some well-marked property; those of most metals end in *um.* *Compounds* are named from their constituent atoms and the names express as fully as possible both the qualitative and quantitative composition of their molecules. Many compounds have beside this *scientific* name, other names; arbitrary, commercial or mineralogical.

13. Chemical Reactions. In general any chemical change is called a *chemical reaction.* The atoms may be changed in kind, in number or in their relations to each other. Substances between which a reaction occurs are called *reagents*. The rearrangement of atoms which is involved in every chemical reaction may be represented by equations whose terms are molecular formulas.

Reactions may be classed as:—

a. *Analytical;* in which complex molecules are resolved into simpler ones.

b. *Synthetical;* in which two or more molecules or atoms unite to form a more complex substance.

c. *Metathetical;* in which an exchange of atoms takes place between molecules.

14. Conditions influencing chemical reaction.

a. Since all chemical reactions take place between molecules and atoms they are favored by conditions which promote molecular freedom and intimate contact of the reagents. Hence solution, and the liquid and gaseous states are especially favorable to atomic rearrangement.

b. If the reagents are in solution, the *insolubility* of a possible compound usually determines its formation. *Precipitation.*

c. The *volatility* of a possible compound at the temperature of the experiment usually determines its formation. *Effervescence.*

d. An element exhibits increased chemical activity at the instant of its liberation from a compound—in the *nascent state*—before the atoms have united to form elementary molecules; for the affinities are unsaturated and no resolution of the molecule must precede the reaction.

15. Relation of chemical affinity to the phys-

ical forces. Heat, light and electricity are powerful agents in producing chemical action.

Substances which have the power of chemical combination with each other possess *chemical energy.* This energy is *potential,* and in the act of chemical union is transformed into various forms of *kinetic energy,* such as light, heat or electricity.

16. **Physical properties important for characterization.** a. *State of aggregation* and the conditions under which change of state occurs;—melting and boiling points; temperature of solidification; conditions for the liquefaction of gases.

b. *Color, odor, taste.*

c. *Specific gravity.* The specific gravity of a *solid* or *liquid* is the ratio of its weight to that of an equal volume of water at 4° C.; or its weight in grams divided by its volume in cubic centimeters. The specific gravity of a *gas* is the ratio of its weight to that of an equal volume of air or of hydrogen weighed under the same conditions of temperature and pressure. In chemistry hydrogen is used as the standard for the specific gravity of gases. Since equal volumes of all gases contain the same number of molecules (cf. 9), and the molecular weight of hydrogen is 2, *the specific gravity of gases is one-half the molecular weight.*

d. *Solubility* in various liquids. Solubility of *solids* is, in general, greater as the temperature is higher. On cooling or evaporation of a saturated solution the solid is deposited, usually in crystalline form. Gases dissolve more freely as the temperature is lower, and are driven out of their solutions by heat.

Further in the case of solids:—

e. *Lustre, hardness, malleability,* &c.

f. *Crystalline form.*

17. **Classification of the Elements.** **a.** The elements are ordinarily divided according to their physical and certain chemical properties into two groups:—*Metals* and *Non-Metals.* It is, however, impossible to draw a sharp line of classification between the groups, and the division is hence somewhat arbitrary and retained chiefly for convenience.

b. The elements may be arranged according to their electro-chemical behavior, as *positive* and *negative.* The non-metals as a class are negative with reference to the metals, and positive or negative towards each other, the relation being entirely a relative one.

c. *The Natural System.* If the elements are arranged in the order of their atomic weights, the series may be broken into several well defined periods, whose members show marked analogies with the corresponding members of the other periods, and a regular gradation of properties among themselves. This may be expressed by saying: The properties of the elements and of their compounds present themselves as *periodic functions* of their atomic weights.

As this system brings elements of similar properties into the same group, it gives an arrangement well adapted to the comparative study of both elements and compounds. (See table on page 79.)

THE NON-METALS.

HYDROGEN.

$H^I=1.$ $H_2=2.$

Occurs *free* in the gases of certain volcanoes and oil wells; also in the atmosphere of the sun and fixed stars; but on the earth is chiefly *combined* in water and in vegetable and animal substances.

Obtained: a. From *water* (H_2O) :—

1. By electrolysis (of acidulated water); hydrogen appears at the negative pole.

2. By the action of some metal: *e.g.* sodium or potassium at ordinary temperature; iron at a red heat.

b. From *acids* by the action of metals.

Prepared usually by the action of zinc on dilute sulphuric acid :—

$$Zn+H_2SO_4=ZnSO_4+H_2.$$

Properties. *Physical.*—When pure, hydrogen is a colorless, odorless and tasteless gas. It is the lightest substance known, being 14.44 times lighter than air, and is consequently chosen as the standard of specific gravity of gases, and its atom is taken as the unit of atomic and molecular weight. At 0°, and under a pressure equal to that of 760 mm. of mercury (*normal conditions* for the measurement of the volumes of gases), one litre weighs 0.0896 (0.089578) gram. Very slightly soluble in water. Has been liquefied by means of intense cold and great pressure.

Chemical.—Combustible, burning with a pale, but very hot flame. Does not support combustion or respiration. It is *positive* in its chemical behavior towards the non-metals and shows other analogies to the metals. Is a constituent of many of the most important compounds.

Used to fill balloons; as fuel in the arts; as *reducing agent* in the laboratory.

History. Apparently known to Paracelsus in the 16th century; but first accurately described by Cavendish in 1766, who called it inflammable air. Lavoisier gave it the name of Hydrogen (*ὕδωρ* and *γεννάω*).

THE HALOGEN GROUP.

Fluorine, Chlorine, Bromine and Iodine form a well characterized natural group. Called Halogens (*ἅλς* and *γεννάω*) because they yield salt-like compounds by direct union with metals. They are all negative monads.

CHLORINE.

$Cl^{I}=35.4$. $Cl_2=70.8$.

Occurs only in combination with metals (chlorides); chiefly with sodium in common salt which exists abundantly in the sea, salt lakes, springs and saline deposits.

Prepared: a. From common salt (NaCl) by the action of manganese dioxide (MnO_2) and sulphuric acid, the reaction comprising two phases:—

1. $2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl$.

2. $2HCl + MnO_2 + H_2SO_4 = MnSO_4 + 2H_2O + Cl_2$.

b. From a solution of hydrogen chloride (HCl) —which has previously been obtained by the reaction indicated in a. 1.—by the action of MnO_2 alone:—

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2.$$

Properties. *Physical.*—Yellowish-green gas of disagreeable and irritating odor. Its specific gravity is 35.4 and it is consequently 2.45 times heavier than air. Can be condensed to a yellow liquid by cold (—40°), or pressure (4 atmospheres at 15°). Somewhat soluble in water (2 vol. are absorbed by 1 vol. water at 20°). By saturating water at 0° with chlorine a crystalline hydrate is obtained, $Cl_2 + 10H_2O$.

Chemical.—Incombustible, but supports the combustion of many bodies. It acts energetically upon most substances; combines directly with all the elements except oxygen, nitrogen, carbon and fluorine; with many of them at ordinary temperatures: *e. g.* phosphorus, antimony, hydrogen (in light). It is a powerful bleaching and disinfecting agent. We may classify the reactions with chlorine under three heads:—

a. Direct union with elements, forming *chlorides.*

b. Substitution; where an atom of chlorine takes the place of an atom of hydrogen—usually in an organic compound:

$$C_xH_yH + Cl_2 = C_xH_yCl + HCl.$$

c. Withdrawal of hydrogen from compounds without replacement; action on a burning candle, turpentine; on water in sunlight or in the presence of a substance which is readily oxidized: $H_2O + Cl_2 = 2HCl + O$. The oxygen is set free in this reaction in the *nascent state* and shows the greatest chemical activity; (indirect oxidation).

The *bleaching and disinfecting power* of chlorine is explained by (b) and (c). The last reaction shows why the presence of moisture promotes the bleaching. Mineral colors in general and carbon (printers' ink) are unaffected by chlorine.

Used as a bleaching and disinfecting agent (chloride of lime); and as reagent in the laboratory, chiefly in aqueous solution, *chlorine water*.

History. First obtained by Scheele, 1774, who called it dephlogisticated marine acid gas; afterwards oxymuriatic acid. Its elementary character was established by Davy, 1810, who named it chlorine (χλωρός).

HYDROGEN CHLORIDE. $HCl = 36.4$.

Occurs only in some volcanic gases.

Formed: a. By union of equal volumes of hydrogen and chlorine; this is effected gradually in diffused light; explosively in direct sun-light, or by a flame: $H_2 + Cl_2 = 2HCl$. The volume of HCl formed equals the sum of the volumes of the hydrogen and chlorine.

b. By the action of chlorine on hydrogen compounds (see above b and c).

c. By the decomposition of chlorides by acids.

Prepared always by the action of sulphuric acid on common salt (see page 12, a. 1). Cf. 14. c.

Properties. *Physical.*—Colorless gas of exceedingly pungent and suffocating odor and acid taste. Specific gravity, 18.2. Condensed by pressure (40 atmospheres at 10°) to a colorless liquid. Remarkably soluble in water (500 : 1 at 0° ; 450 : 1 at 15°) with evolution of heat, and cannot be entirely driven off by boiling. Both the gas and its aqueous solution are called *hydrochloric acid.* On heating, this acid solution if concentrated loses HCl and becomes weaker ; if weak, loses water and becomes stronger ; until finally in each case a solution containing 20.5 per cent. HCl is produced which boils at 110°. On account of its great avidity for water, the gas fumes strongly in moist air.

Chemical.—Both gas and solution are strong *acids,* acting energetically on many metals and their compounds.

Acids are compounds which have a sour or acid taste, redden blue litmus paper, and neutralize bases with the formation of salts. They all contain hydrogen, which can be replaced by a metal.

A Base is a metallic hydroxide which by reaction with an acid forms salts. Soluble bases (alkalies) have a caustic taste and turn red litmus paper blue.

A Salt is a compound formed by the replacement of the whole or part of the hydrogen of an acid by a metal. The salts of hydrochloric acid are *chlorides.*

Uses. Hydrochloric or "muriatic" acid is produced in immense quantities as bye-product in the manufacture of soda. It is used for making chlorine and for many minor purposes in the arts, and is an important reagent in the laboratory.

History. Known to the alchemists as spirit of salt. Priestly first obtained the pure gas in 1772, and in 1810 Davy determined its true composition.

BROMINE.

$Br^{I}=79.8$. $Br_2=159.6$.

Occurs only in combination with metals (bromides), chiefly with sodium, potassium and magnesium, whose bromides are found with sodium chloride, though in much smaller quantities.

Prepared from the bromides contained in the mother liquors of salt works by the action of sulphuric acid and manganese dioxide (cf. chlorine).

Properties. *Physical.*—Bromine is a heavy, dark-red, thin liquid. Specific gravity, 3.15. It freezes to a solid at —7.3°. The liquid evaporates rapidly; at 63° boils. Its vapor has an exceedingly strong and disagreeable odor and is very irritating when inhaled. Less soluble in water than chlorine; dissolves readily in chloroform, ether and in carbon disulphide.

Chemical.—Bromine closely resembles chlorine, but is less energetic, and is set free from its compounds with the metals (bromides) by the latter:

$$KBr+Cl=KCl+Br.$$

Free bromine colors starch-paste yellow.

Used in the preparation of bromides which are employed in photography, medicine, manufacture of certain dye-stuffs, &c.

History. Discovered in the water of the Mediterranean by Balard, 1826, who also gave it the name bromine (βρῶμος).

Hydrogen Bromide, HBr, is *formed:* a. By union of bromine vapor and hydrogen by means of heat or platinum sponge. b. By action of bromine on various hydrogen compounds. c. By decomposition of bromides by water or acids.

Prepared by the reaction of bromine with (red) phosphorus and water :

$$5Br+P+4H_2O=H_3PO_4+5HBr.$$

The reaction similar to that for obtaining HCl can not be employed, because

$$H_2SO_4+2HBr=2H_2O+SO_2+Br_2.$$

Properties. Completely analogous to those of hydrogen chloride. Liquefies at −73° and freezes at −87°. Both the gas and its solution in water are called *hydrobromic acid.*

IODINE.

$I^{I}=126.6.$ $I_2=253.2.$

Occurs in iodides, chiefly as sodium iodide accompanying the chlorides and bromides, and in Chili saltpetre.

Preparation. Certain sea-weeds take up iodides from their solution in sea water, and iodine is principally obtained from the sea-weed collected on the rocky coasts of Ireland, Scotland and France. The sea-weed is dried, burned (or carbonized) and the ash ("Kelp" or "Varec") washed in water from which on evaporation certain other salts crystallize out, leaving "mother-liquors" which still contain all the iodides in solution. Iodine is prepared from this solution and from the mother-liquors of Chili saltpetre by the action of sulphuric acid and manganese dioxide (cf. chlorine); or by means of chlorine.

Properties. *Physical.*—Iodine is a grayish-black, opaque, crystalline solid of metallic lustre. Its specific gravity is 4.95. Volatilizes slowly at ordinary temperatures; at 113°–115° it melts, and boils above 250°, evolving a violet* vapor which is the heaviest known. Very slightly soluble in water; dissolves more readily in alcohol (tincture of iodine); very easily in chloroform, ether or carbon disulphide, and also in an aqueous solution of potassium iodide.

Chemical.—Iodine resembles bromine and chlorine, but acts less energetically than either of these elements, and is set free from its compounds (iodides) by them. It colors starch-paste deep blue. This reaction is very delicate and serves as a means for its detection.

Used in medicine, both free and in combination, also in the dye-stuff industry and in photography.

History. Discovered by Courtois, 1812. Named from the color of its vapor (*ἰώδης*).

Hydrogen Iodide, HI, is formed and prepared in the same way as HBr. Its properties are completely analogous to those of HBr. Easily liquefied (by 4 atmospheres at 0°) and solidifies at −51°. Much less stable than HCl and HBr; decomposes at 180° into hydrogen and iodine; is a powerful reducing agent. Its solution is called *hydriodic acid,* whose salts (iodides) are of great importance in photography and medicine.

---

* Iodine vapor is violet only when mixed (as it usually is) with air. When pure it is deep blue.

FLUORINE.

$F^{I}=19$. ($F_2=38$.)

Occurs only in combination; chiefly in fluor-spar (calcium fluoride) and cryolite (sodium-aluminium fluoride). It has recently been obtained in the free condition by electrolysis of liquid hydrofluoric acid contained in a platinum vessel. It is a colorless gas and is the most active of all the elements at ordinary temperatures, attacking almost all substances with great intensity.

Hydrogen Fluoride, HF, is *prepared* from calcium fluoride by the action of sulphuric acid:—

$$CaF_2+H_2SO_4=CaSO_4+2HF.$$

Properties.—As usually prepared hydrogen fluoride is a colorless gas of extremely irritating and acid odor, which fumes strongly in moist air. Dissolves freely in water and both gas and solution are called *hydrofluoric acid.* It can also be obtained as an anhydrous liquid which boils at 19.4°.

Chemical.—Very active, dissolving most metals with the formation of *fluorides.* It is distinguished from all other substances by its power of dissolving silica, and is extensively used for etching glass.

The gradation of physical properties in the group of halogens is seen by the following table:

	Fluorine.	Chlorine.	Bromine.	Iodine.
At. wts.,	19	35.4	79.8	126.6
Sp. gr. (gas),	—	2.45	5.54	8.7
Melting point,	—	—75°	24°.5	113°
Boiling point,	—	—33.6°	63°	250°
State,	gas	gas	liquid	solid
Color,	colorless	greenish	brown	black

A corresponding gradation in chemical properties is observed—from the extremely energetic fluorine to the comparatively inert iodine.

THE OXYGEN GROUP.

This group includes four analogous elements: Oxygen, Sulphur, Selenium and Tellurium. They are all negative dyads.

OXYGEN.

$O^{II}=16$. $O_2=32$.

Occurrence. Oxygen is the most abundant and widely distributed element. It occurs *free,* but mixed with nitrogen in the air. In *combination* it exists in a large number of substances, both organic and inorganic. It forms $\frac{8}{9}$ of water and about $\frac{1}{2}$ of the solid crust of the earth.

Obtained: a. From the air only by indirect means. b. From acidulated water by electrolysis; appearing at the positive pole. c. From metallic oxides or salts, usually by means of heat; *e. g.*: $2HgO + heat = 2Hg + O_2$.

Prepared by heating potassium chlorate ($KClO_3$), which is usually mixed with manganese dioxide, as in this case the decomposition occurs at a lower temperature and more regularly.

$$2KClO_3 = 2KCl + 3O_2.$$

Properties. *Physical.*—Colorless, odorless, tasteless gas. Slightly heavier than air. Can be condensed to a liquid whose boiling point is —181.4°. Slightly soluble in water (4.1 : 100 at 0°).

Chemical.—Incombustible, but a powerful supporter of combustion. Combines with all the elements except fluorine and acts readily on many compounds; *oxidation and oxides.*

Used to increase the energy of combustion for the production of heat or light. Diluted with nitrogen in the air, it supports respiration and combustion, and plays an important part in most natural chemical changes.

History. Discovered by Priestly, 1774, who called it dephlogisticated air; independently by Scheele, 1775, who called it empyreal air. Named oxygen (*οξύς* and *γεννάω*) by Lavoisier, 1781.

Combustion. A chemical union attended by evolution of light and heat is called a *combustion.* Substances are commonly classed as combustible or incombustible according to their behavior in air (oxygen) which is naturally taken as the typical supporter of combustion. But these terms are readily seen to be entirely relative ones. Thus a jet of air burns in hydrogen just as one of hydrogen does in air; and in general, in the case of gases, the rôles of combustibility and of supporting combustion can be exchanged.

Kindling Temperature. The temperature at which ignition occurs varies widely with different substances; *e. g.* liquid hydrogen phosphide ignites spontaneously at the ordinary temperature; phosphorus takes fire at 50°; coal-gas at a temperature above that of red-hot iron.

Many substances undergo oxidation at temperatures far below that at which they ignite, and in some cases this slow combustion raises the temperature of the body to the ignition point and spontaneous combustion takes place; *e. g.* phosphorus, finely divided iron, oily rags, &c.

OZONE. $O_3=48$.

This *allotropic* modification of oxygen is formed by a condensation of the latter which takes place under various conditions; three volumes of oxygen yielding two volumes of ozone.

$$3O_2=2O_3.$$

In no case is all of the oxygen converted into ozone; only 5 or 6 per cent. being changed under the most favorable conditions. The usual methods for the preparation of ozone are:—1. Slow oxidation of phosphorus in moist air. 2. Silent electrical discharge in air or oxygen.

Ozone occurs in small quantity in the air.

Properties. Ozone has an intense and peculiar odor. Hence its name (ὄζω). It is converted into oxygen when heated to 300°. It can be condensed to a blue liquid.

It is a powerful oxidizing agent; bleaches strongly, attacking all organic substances; destroys the lustre and mobility of mercury; sets iodine free from KI (detection by starch papers moistened with KI solution). Ozone is absorbed by turpentine. It has been used for bleaching engravings discolored by age.

COMPOUNDS OF OXYGEN WITH HYDROGEN.

Two compounds of oxygen and hydrogen are known: Hydrogen monoxide and Hydrogen dioxide.

HYDROGEN MONOXIDE. $H_2O=18$.

Occurs abundantly in nature in all three states of aggregation. a. As gas or "aqueous vapor" in the atmosphere. b. As liquid—water—in the ocean, in rivers, springs, &c.; in all animals and plants. c. As solid—snow and ice. It is found further in many minerals and salts, often as a constituent essential to their crystalline existence; *water of crystallization. Efflorescence* and *deliquescence.*

Formed: a. By union of two volumes of hydrogen and one volume of oxygen:

$$2H_2+O_2=2H_2O.$$

The union may be effected by a flame, electric spark or finely divided (spongy) platinum. The reaction is a very energetic one as shown by the detonation when the mixed gases are ignited, and the intense heat which attends the combustion of hydrogen in air or oxygen. *Oxy-hydrogen blowpipe*

b. Whenever substances containing hydrogen are burned or slowly oxidized; hence in the combustion of all organic substances, in respiration, &c.

c. In many chemical reactions.

Properties. *Physical.*—Pure water is a tasteless, odorless liquid, colorless in small quantity, but showing a bluish-green tinge when light passes through a considerable thickness. It is densest at 4°, expanding when heated above or cooled below this temperature. At 0° it *usually* freezes to ice, which always melts at 0°. This change is accompanied by a considerable change of volume—100 vol. water at 0° yielding 107 vol. ice at 0°. At 100° it boils. Water at 4° is taken as the standard of specific gravity for solids and liquids, and the unit of weight (gram) is the weight of one cubic centimeter of water at 4°. Water has a greater capacity for heat than any other substance, and hence the amount of heat necessary to raise a unit weight of water through 1° is taken as the unit of heat. Water dissolves very many solids, liquids, gases, and hence is never found pure in nature. The purest natural water is rain or snow water. Water containing certain salts—notably calcium carbonate (held in solution by carbonic acid) and sulphate—is called *hard.* Hardness due to the presence of calcium carbonate is called *temporary,* because it can be removed by boiling or by addition of milk of lime; if due to sulphates it is *permanent.* Water is *purified* (a) from suspended particles by filtration or subsidence; (b) from dissolved substances by precipitation or distillation; (c) from gases by continued boiling.

Chemical.—Hydrogen monoxide is a neutral substance, neither acid nor alkaline in character. It is decomposed into oxygen and hydrogen at a very high temperature or by a powerful electric current; it oxidizes certain metals (*e. g.* potassium) at ordinary temperatures; others, only when heated (*e. g.* iron)—hydrogen being set free at the same time (cf. p. 10). Is decomposed by chlorine in sunlight (cf. p. 13. c).

History. The composition of water was first determined by Cavendish, 1781.

Hydrogen Dioxide, H_2O_2, is the result of the action of dilute acids on certain peroxides:—

$$BaO_2 + H_2SO_4 = BaSO_4 + H_2O_2.$$

Properties.—In its most concentrated form, hydrogen dioxide or peroxide of hydrogen is a colorless syrupy liquid of sp. gr. 1.45, which does not solidify at —30°; unstable, decomposing easily with liberation of oxygen. In dilute solution in water more stable. It is a powerful oxidizing agent, acting like ozone with which it has strong analogies. It is detected by the blue color it produces in an acid solution of potassium chromate.

Uses.—Hydrogen is manufactured in considerable amounts for use in bleaching and medicine.

Hydroxyl, $(HO)^{I}$. A *group* which often plays the part of a monad element. Incapable of isolation, it is believed to be combined with itself in hydrogen dioxide—HO–OH.

COMPOUNDS OF OXYGEN WITH THE HALOGENS AND HYDROGEN.

All the halogens except fluorine combine with oxygen to form oxides, or with oxygen and hydrogen to form acids.

OXIDES and OXY-ACIDS OF CHLORINE.

Oxides.		Acids.	
Cl_2O	*monoxide.*	$HClO$	*hypochlorous.*
Cl_2O_3	*trioxide.*	$HClO_2$	*chlorous.*
ClO_2	*dioxide.*	——	
——		$HClO_3$	*chloric.*
——		$HClO_4$	*perchloric.*

The oxides are heavy yellowish gases, very unstable and acting as powerful oxidizing and bleaching agents. Dissolved in water they form hypochlorous, chlorous and a mixture of chlorous and chloric acids respectively.

The acids are also unstable compounds forming salts by replacement of the hydrogen by a metal. The most important salts of this series of acids are: The *hypochlorites* (calcium hypochlorite, bleaching powder) and the *chlorates* (potassium chlorate).

OXY-ACIDS OF BROMINE. $HBrO$, *Hypobromous acid.* $HBrO_3$, *Bromic acid.* $HBrO_4$, *Perbromic acid.*

These acids are completely analogous to the corresponding chlorine compounds. No oxide of bromine is known.

OXIDES AND OXY-ACIDS OF IODINE.

I_2O_5, *Pentoxide.* HIO_3, *Iodic acid.* HIO_4, *Periodic acid.* These compounds are similar to the corresponding chlorine and bromine compounds, but are much more stable.

SULPHUR.

$S^{II}=32$. $S_2=64$ (above 1000°) ; $S_6=192$ (at 500°).

Occurs *free* (native) in volcanic regions (Sicily) on the surface and imbedded in marl, limestone, &c., in vast deposits (sulphur mines). *Combined* (a) with metals in sulphides—pyrites, galena, &c.; (b) with metals and oxygen in sulphates—gypsum, &c.; (c) in certain organic compounds, and in "sulphur springs."

Extracted from the volcanic ore by melting it out; or from iron pyrites FeS_2 by distillation. The crude sulphur is purified by distillation and yields *flowers of sulphur* when the vapor is suddenly cooled; or *roll sulphur* when the melted distillate is cast into sticks.

Properties. *Physical.*—Sulphur is a yellow brittle solid in its ordinary condition; but is capable of existing in at least three allotropic forms.

a. *Ordinary rhombic sulphur.* Pale yellow, hard and very brittle, a non-conductor of heat and electricity. (Crackles and cracks under the heat of the hand.) Sp. gr. 2.07. Melts at 114°.5 to a thin amber-colored liquid, which on further heating becomes darker and thicker until at 250° it is nearly black and so viscid that it cannot be poured. Above 300° it grows thin again and boils at about 450° yielding a deep red vapor. Insoluble in water; dissolves readily in sulphur chloride and in carbon disulphide.

b. *Monoclinic sulphur* is formed when melted sulphur cools, in long elastic amber-colored transparent needles. Specific gravity, 1.96. Melts at 120°. Soluble in the same solvents as (a). This modification is unstable. The crystals soon become opaque and brittle, and the sulphur is found to have changed into (a).

c. *Plastic sulphur* is obtained as a yellowish brown, soft, tenacious mass when melted sulphur at about 300° is poured into cold water. It soon becomes hard and brittle from conversion into (a). Carbon disulphide dissolves only a part of it, leaving *amorphous insoluble sulphur.*

Sulphur may be separated from solutions of certain of its compounds in a very finely divided condition, giving to the liquid the appearance of milk; *e. g.* on adding HCl to a solution of sodium thiosulphate, $Na_2S_2O_3$.

The specific gravity of sulphur vapor varies with the temperature in such a way as to indicate that at about 500° the molecule consists of six atoms, while at 1000° and above it consists of two atoms.

Chemical.—Sulphur burns when heated to 260° with a pale blue flame. It unites directly with very many elements forming *sulphides* which show strong analogies with the oxides.

Uses. Sulphur is used as an ingredient of gunpowder and fire-works, in "vulcanizing" india rubber, in the straw and woollen bleachery, in making sulphuric acid and for some minor purposes.

History. Known from the earliest times—called Brimstone or Brennestone, and believed by the alchemists to be the principle of combustibility.

COMPOUNDS OF SULPHUR AND HYDROGEN.

Two compounds of sulphur and hydrogen are known: Hydrogen monosulphide and hydrogen persulphide.

HYDROGEN MONOSULPHIDE.

$H_2S=34$.

Occurs in volcanic gases and sulphur springs.

Formed from various organic substances by decay or by destructive distillation; by heating various hydrocarbons (*e. g.* paraffine) with S; by leading hydrogen over heated sulphides or through boiling sulphur; and by the decomposition of sulphides by acids.

Prepared by the action of hydrochloric or sulphuric acid upon iron sulphide:

$$FeS+2HCl=FeCl_2+H_2S.$$

Properties. *Physical.*—Colorless, poisonous gas of very offensive odor. Can be condensed to a colorless liquid, which solidifies at —85°. Water dissolves 3 to 4 times its own volume of the gas.

Chemical.—Hydrogen sulphide burns when ignited in air or oxygen to H_2O and SO_2. If the supply of oxygen is insufficient (less than 1½ times the volume of the H_2S), or if a cold body is introduced into the flame, the combustion is incomplete and sulphur is deposited. In aqueous solution it is decomposed by the oxygen of the air with deposit of sulphur. The halogens decompose it in a similar way.

It is a reducing agent. Attacks certain metals (*e. g.* silver) at ordinary temperatures and nearly all metals when heated, with the formation of metallic sulphides and liberation of hydrogen; precipitates many metals from their solutions as sulphides. Its presence is detected by the odor or by the blackening of paper saturated with a lead solution.

Uses. Hydrogen sulphide or sulphuretted hydrogen is an important laboratory reagent, being employed both as gas and in solution.

Hydrogen Persulphide, H_2S_2. A heavy yellowish oil which decomposes readily into hydrogen sulphide and sulphur. It is analogous to hydrogen peroxide.

SULPHUR AND THE HALOGENS.

Three compounds of sulphur and chlorine are known: S_2Cl_2, SCl_2 and SCl_4.

Sulphur Monochloride, $S_2Cl_2=134.8$, is the most stable of the chlorides of sulphur. It is obtained as an amber liquid of unpleasant penetrating odor by leading dry chlorine over heated sulphur. It is one of the best solvents of sulphur known and employed on this account in the process of vulcanizing rubber. It boils at 138° and decomposes with water:

$$2S_2Cl_2+2H_2O=4HCl+SO_2+3S.$$

Sulphur Dichloride, SCl_2, a reddish brown liquid, which decomposes when heated into chlorine and the monochloride; obtained by saturating sulphur monochloride with chlorine at ordinary temperature.

Sulphur Tetrachloride, SCl_4, formed by saturating the dichloride with chlorine at $-22°$. A yellowish brown liquid which decomposes into chlorine and the dichloride when taken out of the freezing mixture.

One compound of Sulphur with *bromine,* S_2Br_2; and two with *iodine,* S_2I_2 and SI_6, are known.

SULPHUR, OXYGEN AND HYDROGEN.

Sulphur forms two important oxides, SO_2 and SO_3; which are the *anhydrides* of the acids, H_2SO_3 and H_2SO_4.

SULPHUR DIOXIDE. $SO_2=64$.

Occurs in volcanic gases and in certain springs.

Formed by burning sulphur or sulphides in air or oxygen; by heating metallic oxides with sulphur; by decomposing sulphites with an acid; heating certain sulphates; from concentrated sulphuric acid by the action of sulphur, charcoal or certain metals.

Prepared: a. For technical purposes by burning sulphur or roasting pyrites:

$$2FeS_2+11O=Fe_2O_3+4SO_2.$$

b. In the laboratory by the action of copper on sulphuric acid:

$$2H_2SO_4+Cu=CuSO_4+2H_2O+SO_2.$$

Properties. *Physical.*—Colorless gas of suffocating odor. Condenses under a pressure of 2 to 3 at. or when cooled to —15° to a colorless liquid which solidifies at —76° and boils at —10.5°, producing great cold by its rapid evaporation. The gas dissolves freely in water (40 : 1 at 20°), forming *sulphurous acid,* H_2SO_3.

Chemical.—Incombustible and a non-supporter of combustion, but unites with oxygen in the presence of platinum sponge to form SO_3. In the presence of water (as sulphurous acid) it is an energetic reducing agent, withdrawing oxygen from compounds and becoming H_2SO_4. This behavior renders it an important antiseptic, disinfectant and bleaching agent. Cf. chlorine. With H_2S:

$$SO_2+2H_2S=2H_2O+3S.$$

Uses. Sulphur dioxide is used for bleaching straw, woolen and silk ; to prevent fermentation ; as an antichlor in the paper manufacture, and in the manufacture of sulphuric acid.

Sulphurous Acid. H_2SO_3.

Known only in aqueous solution, as this separates when heated into water and sulphur dioxide. It contains two hydrogen atoms replaceable by metals, and is hence a *bibasic* acid, forming *acid* or *normal sulphites* according as one or both atoms of hydrogen are replaced.

Sulphur Trioxide, SO_3, is *formed* when a mixture of sulphur dioxide and oxygen is passed over heated platinum sponge ; by distillation of fuming (Nordhausen) sulphuric acid ; or by heating concentrated sulphuric acid with phosphorus pentoxide.

Properties.—White silky crystals. Melts at 16°, boils at 46°, and changes when kept melted below 25° into a solid which melts at 50° and returns to the first modification.

It unites with water with evolution of great heat, forming H_2SO_4 (hence called sulphuric anhydride).

SULPHURIC ACID. $(HO)_2SO_2=98$.

Occurs *free* in minute quantities in the waters of certain springs; *combined* in sulphates especially as calcium sulphate (gypsum).

Formed by oxidation of sulphur (by concentrated nitric acid), sulphur dioxide or other compounds of sulphur in the presence of water, or by direct union of sulphur trioxide and water. It cannot be formed from sulphates by the action of acids (as HCl is from chlorides).

Prepared by oxidation of sulphur dioxide in the presence of water by means of certain oxides of nitrogen :—

1. a. $S+O_2=SO_2$.
Or b. $2FeS_2+11O=Fe_2O_3+4SO_2$.
2. $SO_2+H_2O+NO_2=H_2SO_4+NO$.

The NO thus formed is reconverted by the oxygen of the air into NO_2 which again enters into reaction with fresh molecules of SO_2 and H_2O.

Properties. The acid that is produced in the leaden chambers by the above reaction contains about 60 per cent. of H_2SO_4, and has a specific gravity of 1.5. By evaporation in leaden and platinum vessels, the specific gravity is raised to 1.80–1.84 and the percentage of H_2SO_4 to 90–98. This is the commercial *oil of vitriol.* Pure H_2SO_4 has the specific gravity 1.842 (at 12°). On cooling it, crystals are obtained, which melt at 10°.5. Sulphuric acid is a colorless, oily and very corrosive liquid. It has a great affinity for water, uniting with it with evolution of great heat; absorbs moisture from the air and is used for drying air and other gases; and even withdraws the elements of water from organic substances brought in contact with it, completely charring them.

The acid is bibasic and forms both acid and normal salts—*sulphates.*

Uses. Sulphuric acid is the most important substance used in chemical manufactures.

OTHER ACIDS OF SULPHUR.

Besides sulphurous and sulphuric acids, some seven sulphur acids are known, of which the most important is the fuming, Nordhausen or *pyrosulphuric* acid $(HO)_2O(SO_2)_2$; *prepared* by the distillation of "green vitriol" (ferrous sulphate); or by treating ordinary sulphuric acid with SO_3. It is a fuming, oily liquid, more or less dark colored from organic matter with which it has come in contact and charred. On cooling, it yields colorless crystals of $H_2S_2O_7$ which melt at 35°. Decomposes on heating into H_2SO_4 and SO_3; water converts it into H_2SO_4.

History. Sulphuric acid was prepared in the 15th century from green vitriol. The present method was first employed in England about the middle of the last century.

SELENIUM.

Se=78.8.

Occurs widely distributed but in very small quantities and always in combination. Discovered in 1817 by Berzelius, in the lead chamber deposits of the sulphuric acid works at Gripsholm. Named from *σελήνη*, the moon, from analogy with tellurium.

Properties. Selenium is known in two allotropic modifications, and is in other respects the complete analogue of sulphur.

Melts at 217°. The crystalline modification conducts electricity and its conductivity is increased by exposure to light.

The compounds of selenium are analogous to those of sulphur.

TELLURIUM.

Te=125.

Occurs rarely, both native and combined with certain metals. Identified as a new substance by Klaproth, 1798; named from *tellus*.

It is a silver white brittle solid of metallic lustre. Melts at 455°. Physically it is a metal; chemically it resembles sulphur and selenium.

COMPARISON OF THE ELEMENTS OF THE OXYGEN GROUP.

	Oxygen.	Sulphur.	Selenium.	Tellurium.
Atomic wt.,	16	32	78.8	128
State,	gas	solid	solid	solid
Spec. grav.,	—	1.96–2.07	4.5–4.8	6.18
Melting point,	—	114°.5	217°	455°
Boiling point,	—181°	450°	666°	white heat.

The chemical affinity for hydrogen and chlorine decreases from oxygen to tellurium.

NITROGEN GROUP.

To this group belong Nitrogen, Phosphorus, Arsenic, Antimony and Bismuth.

NITROGEN.

$N^{III \text{ and } V}=14$. $N_2=28$.

Occurs *free* but mixed with oxygen in the air. *Combined* in many organic compounds (*e. g.* albumen), in ammonia, nitric acid and nitrates (*e. g.* KNO_3).

Prepared: a. From air by removing the oxygen by means of an oxidizable body, *e. g.* phosphorus, copper, iron.

b. From certain compounds: by the action of chlorine upon an *excess* of aqueous ammonia; from ammonium nitrite by heat:

$$NH_4NO_2=2H_2O+N_2.$$

Properties. *Physical.*—Colorless, odorless, tasteless gas. Somewhat lighter than air. Has been condensed to a liquid by great cold and pressure. At a pressure of 4 mm. liquid nitrogen solidifies completely, giving a temperature of —225°, which is the lowest that has been obtained. Slightly soluble in water (1:50).

Chemical.—Incombustible and a non-supporter of combustion; in the free state very inert, uniting directly with only a few substances and with these only at a high heat. It is, however, an essential constituent of some of the most energetic compounds known: ammonia, nitric acid, nitro-glycerine, prussic acid and the alkaloids.

History. Discovered 1772, by Rutherford and by Priestly. Scheele recognized it as a constituent of the air. Lavoisier was the first to consider it an element and named it azote (*a* and *ζωή*)—a name still retained by the French. Chaptal suggested the name nitrogen (Nitrum and *γένναω*) with reference to its presence in nitre.

THE ATMOSPHERE.

The existence of the invisible ocean of air at the bottom of which we live is proved: a. By the resistance it offers to motion and the effects produced by its motion (winds). b. By its weight and consequent pressure. The first distinct proof of the pressure of the atmosphere is due to Torricelli, 1643. Measurement of pressure by the *Barometer*. *Normal pressure,* 760 mm. of mercury, equal to 1033.3 grams on every square centimeter (or about 15 pounds on every square inch). One litre of pure dry air at 0° and 760 mm. weighs 1.2931 grams (about $\frac{1}{773}$ of the weight of water).

Composition. Air consists chiefly of nitrogen and oxygen in about the following proportions:

	By Volume.	By Weight.
Nitrogen,	79	77
Oxygen,	21	23
	100	100

That these gases are simply mixed and not in chemical combination is proved by the following facts:—

1. The elements are not present in the proportion which their atomic weights demand for a chemical compound, and this proportion varies (within certain small limits).

2. Air may be artificially prepared by mixing oxygen and nitrogen in the proper proportion and its formation is unattended by any evidence of chemical union: evolution of heat or contraction of volume.

3. The proportions of nitrogen and oxygen are changed by simply dissolving air in water; *i. e.* each gas dissolves independently and this it could only do when free.

	Air Dissolved in Water.	Air Undissolved in Water.
Nitrogen,	65.08	79
Oxygen,	34.92	21

On account of the inert character of nitrogen, we may regard air as diluted oxygen.

Beside these chief constituents, the air always contains aqueous vapor, carbon dioxide, ammonia and ozone, and often minute quantities of other gases and dust (organic and inorganic) as impurities.

The proportion of aqueous vapor which air can contain depends upon the temperature, being larger as the temperature is higher. The ratio of the amount actually present to that which the air could contain, expressed as a percentage, is called the *humidity* of air. Precipitation of aqueous vapor.

The average amount of carbon dioxide in country air is about 4 parts in 10,000. More is found in towns and less over the sea. This gas serves as the chief food of plants, which utilize its carbon and part of its oxygen, giving back the rest of the oxygen to the air.

By the processes of combustion (lamps, &c.) and respiration, part of the oxygen of the air is replaced by carbon dioxide. Air containing four per cent. of carbon dioxide puts out a candle, but supports respiration for a short time. A slightly larger amount destroys animal life at once by suffocation. The oppressiveness and ill effects of close rooms is due not so much to carbon dioxide as to the presence of decomposing organic matters which are given off with it by the lungs. The amount of carbon dioxide serves however as an index of the contamination of the air. (Ventilation.)

The other substances are present in variable and extremely minute quantities.

AMMONIA. $NH_3=17$.

Occurs combined with acids in the air, in natural waters and in the earth, but only in very small quantities.

Formed: a. In small quantities by the direct union of nitrogen and hydrogen under the silent electric discharge.

b. By the action of nascent hydrogen upon oxygen compounds of nitrogen (*e. g.* when zinc is dissolved in nitric acid).

c. By the decay, putrefaction or dry distillation of nitrogenous organic bodies.

d. By the action of an alkali (lime) on salts of ammonium.

Preparation. The ammoniacal liquor which is obtained as a bye-product in the manufacture of coal gas serves as the source of the chloride or sulphate of ammonium, which is purified by sublimation or crystallization and then heated with lime :—

$$2NH_4Cl + CaO = CaCl_2 + H_2O + 2NH_3.$$

Properties. *Physical.*—Ammonia is a colorless gas of very pungent and suffocating odor, about three-fifths as heavy as air. Condensed by pressure (6–7 at. at 15°), or cold (—40°) to a colorless liquid, which solidifies at —80°. The gas is exceedingly soluble in water (1050 : 1 at 0° ; 730 : 1 at 15°) with evolution of heat, forming a solution known as *ammonium hydroxide.* All the gas can be driven off by heat.

Chemical.—Ammonia is a non-supporter of combustion and does not burn under ordinary conditions in air. In oxygen, however, it burns with a pale yellow flame, and can be oxidized without flame in air or oxygen by the influence of a heated spiral of platinum wire. Both gas and aqueous solution are strongly *basic.* They turn red litmus paper blue and with acids yield *ammonium* salts. The gas unites directly with acid vapors, forming dense clouds of the finely divided salt: *e. g.* $NH_3 + HCl = NH_4Cl$.

In these salts the monad group NH_4 acts as a metal. The gas is decomposed by continued passage of the electric spark into nitrogen and hydrogen which occupy double the volume of the ammonia:

$$2NH_3 = N_2 + 3H_2.$$

Uses. It is employed in making artificial ice; in the manufacture of soda, &c., and its solution is an important reagent in the laboratory.

History. Known to the alchemists and, until 1774, only in aqueous solution—"spirits of hartshorn." The gas was discovered by Priestly, 1774, and called *alkaline air*. The name ammonia is a revival from the old name *sal-armoniacum*, which was given to the chloride of ammonium.

NITROGEN AND THE HALOGENS.

Nitrogen Chloride, NCl_3, formed by the action of chlorine on ammonium chloride (and hence by the action of chlorine in excess on ammonium hydroxide), is a yellow, oily liquid, exceedingly explosive and dangerous.

A similar bromine compound is known.

Nitrogen Iodide, NI_3, is prepared by pouring concentrated ammonia on powdered iodine. It is a black powder which when dry decomposes with violent detonation if touched or heated.

NITROGEN, OXYGEN AND HYDROGEN.

NITRIC ACID. $HNO_3 = 63$.

Occurs *combined* in its salts, chiefly as nitrates of potassium (nitre or salpeter) and sodium.

Nitrates are formed by the decomposition of nitrogenous organic matter in the presence of alkalies (natural and artificial nitre beds).

Formed: a. From nitrogen, oxygen and water when strong electric sparks are passed through air in the presence of water.

b. From nitrates by the action of sulphuric acid.

Prepared from sodium or potassium nitrate by distillation with concentrated sulphuric acid:

$$2KNO_3 + H_2SO_4 = K_2SO_4 + 2HNO_3.$$

Purified by distillation with an equal volume of concentrated sulphuric acid.

Properties. *Physical.*—Colorless, fuming liquid, specific gravity 1.53. At 86° it begins to boil with partial decomposition into nitrogen peroxide, oxygen and water. If distilled, the boiling point rises and the decomposition continues, until at 120.°5 it ceases, and a 68 per cent. acid of specific gravity 1.42, comes over with this as constant boiling point. An acid of the same strength is also obtained when a weaker acid is distilled.

Chemical.—A strong, highly corrosive acid, which acts as a powerful oxidizing agent. It attacks many metals energetically, dissolving most of them with the formation of *nitrates*. Nitrogenous animal substances—silk, wool, &c.—are colored yellow by it, and certain organic substances—cotton, glycerine, &c.,—are converted into powerful explosives.

Uses. Nitric acid is used in making nitro-glycerine, gun-cotton, silver nitrate and other preparations; it is also employed for etching on metals, and as an important laboratory reagent.

History. Known to Geber in the 8th century. Its composition was determined in 1785 by Cavendish.

Aqua Regia is the name given to a mixture of nitric and hydrochloric acids, frequently used for dissolving the "noble" metals—gold, platinum, &c.—and other difficultly soluble substances. Its solvent power depends upon the evolution of nascent chlorine.

OXIDES OF NITROGEN.

Five oxides of nitrogen are known: N_2O, NO, N_2O_3, NO_2, and N_2O_5. None of them occur natrally.

NITROGEN MONOXIDE. $N_2O=44$.

Formed from nitric acid by the action of reducing agents.

Prepared by heating ammonium nitrate:—

$$NH_4NO_3=2H_2O+N_2O.$$

Properties. *Physical.*—Colorless gas, of sweetish taste and odor. About 1½ times heavier than air. Condenses under a pressure of 30 at. at 0° to a colorless liquid which freezes by spontaneous evaporation to a snow-like solid which melts at —99°. Soluble in water (1.3 : 1 at 0°).

Chemical.—Incombustible, but supports combustion with nearly the same energy as oxygen. It is incapable of supporting respiration. Inhaled in small quantities, it produces a transient exhilaration and is hence called *laughing gas;* in somewhat larger amount it produces insensibility.

Used as an anaesthetic in minor surgical operations, especially in dentistry.

History. Discovered by Priestly, 1772. Davy examined it, 1809, and discovered its exhilarating effect. Wells first used it as an anaesthetic in 1845.

NITRIC OXIDE. $NO=30$.

Formed from nitric acid by the reducing action of the nascent hydrogen produced when certain metals are dissolved in the acid.

Prepared most conveniently by dissolving copper in dilute nitric acid:—

a. $3Cu+6HNO_3=3Cu(NO_3)_2+6H$.

b. $2HNO_3+6H=4H_2O+2NO$.

Properties. *Physical.*—Colorless gas. Has been liquefied by great cold and pressure. Slightly soluble in water; readily soluble in a solution of ferrous salts, which it colors reddish brown and from which it is expelled by heat.

Chemical. — Incombustible. Extinguishes most burning bodies which are brought into it. At a red heat is decomposed into nitrogen and oxygen, and consequently supports combustion when this is already energetically in progress. Mixed with vapor of carbon disulphide, it burns with a brilliant actinic light. Unites directly at ordinary temperature with oxygen (oxygen gas or air) with evolution of heat and formation of reddish fumes consisting of nitrogen trioxide or peroxide, according to the proportions in which the NO and oxygen are present. It also unites directly with chlorine, forming NOCl.

Nitrogen Trioxide, $N_2O_3=76$, is formed by direct union of four volumes of nitric oxide and one volume of oxygen, and by the reducing action of starch or arsenious oxide on nitric acid. It is a gas which readily condenses to a dark blue liquid. Dissolves in cold water forming a dilute solution of *nitrous acid.*

Nitrous Acid, HNO_2, is known only in dilute solution. On warming this solution it decomposes into nitric acid, water and nitrogen dioxide.

The salts of nitrous acid, the *nitrites* (*e. g.* KNO_2), are stable bodies, formed by the action of the acid on metallic oxides or by reduction of nitrates.

Nitrogen Peroxide, NO_2, is formed by the direct union of nitric oxide and oxygen; by heating nitrates (lead nitrate); or by reduction of nitric acid. It is a liquid, solidifying at —9° in colorless crystals. At 10° the liquid is yellow and at higher temperatures becomes darker and boils at 22°, yielding a reddish brown vapor of very unpleasant odor, consisting of a mixture of NO_2 and N_2O_4. Decomposed by water; a little cold water forming with it nitric and nitrous acids; while, with an excess of water at ordinary temperature, it gives nitric acid and nitric oxide.

Nitrogen Pentoxide, N_2O_5, is formed by withdrawal of the elements of water from nitric acid (by phosphorus pentoxide), or by the action of chlorine on silver nitrate. Colorless crystals, which fuse at 30° to a dark yellow liquid, which boils at about 45° with partial decomposition. Unites with water with much heat forming nitric acid.

PHOSPHORUS.

$P^{\text{III and V}}=31. \qquad P_4=124.$

Occurs only in combination. The most abundant compound is calcium phosphate, which is the chief constituent of the minerals apatite and phosphorite and forms about 58 per cent. of the bones of animals. Phosphates are present in all fertile soils, and form an essential constituent of the food of both plants and animals. In plants they are found especially in the fruit and seed; in animals, in the bones, tissues and nervous matter.

Preparation. Phosphorus is always prepared from calcium phosphate, and usually from bone-ash by the following process :—

(1.) The bone-ash—which contains about 86 per cent. of calcium phosphate—is treated with dilute sulphuric acid, which converts the insoluble normal phosphate into a soluble acid phosphate :—

$$Ca_3(PO_4)_2+2H_2SO_4=2CaSO_4+H_4Ca(PO_4)_2.$$

(2.) The solution of acid calcium phosphate is poured off from the insoluble calcium sulphate formed in the reaction ; concentrated by evaporation ; mixed with charcoal, dried and heated. The acid phosphate is thus converted into calcium metaphosphate :—

$$H_4Ca(PO_4)_2 = 2H_2O + Ca(PO_3)_2.$$

(3.) The mixture of metaphosphate and charcoal is now placed in fire-clay retorts and heated to bright redness: phosphorus distils and is condensed under water, while calcium pyrophosphate is formed in the retorts:—

$$3Ca(PO_3)_2 + 10C = Ca_3(PO_4)_2 + 10CO + 4P.$$

The crude phosphorus is purified by distillation, or by filtration (after melting in hot water) through chamois leather, and cast into sticks.

Properties. Several allotropic modifications of phosphorus are known. a. *Ordinary phosphorus* is a nearly colorless translucent solid. Specific gravity, 1.84. Brittle at 0°, it becomes soft like wax at 15°, and melts at 44°. Heated in an atmosphere free from oxygen it boils at about 290°, yielding a colorless vapor. Volatilizes to a slight extent with steam; (detection in cases of phosphorus poisoning). Insoluble in water; readily soluble in carbon disulphide and crystallizes easily from this solution. Extremely inflammable and always kept under water. In the light it grows yellow and becomes covered with a white opaque crust. In the air it oxidizes slowly at all temperatures and is luminous in the dark. Ignites at 50° and often spontaneously from the heat of the slow oxidation. (Ignition of finely divided phosphorus deposited from solution in carbon disulphide by evaporation.) The slow oxidation is accompanied by white fumes of garlic-like odor which are very poisonous. Phosphorus burns in oxygen or air to P_2O_5; unites directly with the halogens.

b. *Red phosphorus* is obtained by heating common phosphorus to 240–250° in a closed vessel. It is a reddish brown powder. Specific gravity, 2.15. Insoluble in carbon disulphide. Does not oxidize in air or oxygen until heated to 260°, when it is converted into the ordinary modification; has no odor and is not poisonous.

c. *Other modifications*—metallic phosphorus, black phosphorus and white phosphorus—of less importance have been observed.

Treated with oxidizing agents (*e. g.* HNO_3), phosphorus is slowly converted into phosphoric acid.

Uses. Phosphorus is used in the laboratory, in the manufacture of aniline colors, in the preparation of phosphorbronze, as rat poison, &c.; but 90 per cent. of all that is produced is used in making *matches.* For this purpose both the common and the red modification are employed, the latter for the so-called *safety matches.*

History. First prepared by Brande, 1669. Its existence in bones was discovered by Gahn, 1769, and in 1775 Scheele proposed a method for obtaining it from this source. The name phosphorus (φῶς and φέρω) was first given to a sulphide of barium which was luminous in the dark and called *Bonnonian phosphorus.* The true phosphorus was distinguished from this as *phosphorus mirabalis* or *igneous,* and later as *English phosphorus.* The red modification was discovered, 1845, by v. Schrötter. Phosphorus matches appeared in 1833.

PHOSPHORUS AND HYDROGEN.

Three compounds of phosphorus and hydrogen are known: PH_3, P_2H_4 and P_4H_2. None of them occur naturally.

PHOSPHINE. $PH_3=34$.

Prepared: a. By heating phosphorus with a solution of sodium or potassium hydroxide:—

$$3KOH+P_4+3H_2O=3KH_2PO_2+PH_3.$$

b. By the decomposition of calcium phosphide by water or hydrochloric acid.

c. By the decomposition of phosphonium iodide, PH_4I, in water.

Properties. *Physical.*—Colorless, poisonous gas of very disagreeable garlic-like odor, very slightly soluble in water. Spontaneously inflammable as usually prepared, because of the presence of P_2H_4; when pure it ignites only above 100°. It is feebly basic (cf. NH_3) uniting directly with HI or HBr, forming iodide or bromide of *phosphonium* (as the monad group PH_4 is called).

Liquid Hydrogen Phosphide, P_2H_4, separated from impure phosphine by cooling, is a spontaneously inflammable, very unstable liquid, readily splitting up into PH_3, and *solid hydrogen phosphide,* P_4H_2; the latter is a yellow powder igniting at 160°.

PHOSPHORUS AND THE HALOGENS.

Phosphorus forms with the halogens the following compounds:

——	——	——	PI_2
PF_3	PCl_3	PBr_3	PI_3
PF_5	PCl_5	PBr_5	PI_5

All except PF_3 and PF_5 are formed by direct union of the elements, which takes place with great energy. The chlorine compounds will serve as types of all.

Phosphorus Trichloride, PCl_3, is prepared by leading a current of dry chlorine over gently warmed phosphorus. It is a colorless liquid, of pungent odor. Specific gravity, 1.61. Boils at 76°. Fumes in moist air; is decomposed by water into hydrochloric and phosphorous acids: $PCl_3 + 3H_2O = H_3PO_3 + 3HCl$.

Phosphorus Pentachloride, PCl_5, is formed by saturating PCl_3 with chlorine. It is a white crystalline powder, which sublimes at 100° without melting. In moist air it deliquesces, forming phosphorus oxychloride, $POCl_3$ and HCl :—

$$H_2O + PCl_5 = POCl_3 + 2HCl.$$

A similar reaction takes place between PCl_5 and many compounds containing hydroxyl (OH) or oxygen, and serves as a means of introducing chlorine in the place of this element or group.

The $POCl_3$ formed reacts further with water, forming phosphoric and hydrochloric acids:

$$POCl_3 + 4H_2O = (HO)_3PO + 5HCl.$$

PHOSPHORUS AND OXYGEN.

Two oxides of phosphorus are known: P_2O_3 and P_2O_5.

Phosphorus Trioxide, P_2O_3, is produced when phosphorus is heated in a slow current of dry air. It is a white powder of garlic-like odor. Can be readily sublimed. Brought into the air it burns to P_2O_5. It combines with water instantly, forming phosphorous acid:—

$$P_2O_3+3H_2O=2H_3PO_3.$$

Phosphorus Pentoxide, P_2O_5, is formed by burning phosphorus in dry air or oxygen. It is a voluminous white powder, very hygroscopic, deliquescing in moist air with formation of metaphosphoric acid, HPO_3. It is often used in the laboratory as a dessicating agent and to withdraw the elements of water from compounds—*e. g.* in the preparation of N_2O_5. Cf. p. 44.

ACIDS OF PHOSPHORUS.

Hypophosphorous Acid, H_3PO_2. A salt of this acid is formed when a solution of an alkali is warmed with phosphorus (cf. p. 48). The acid may be obtained from the barium salt by means of sulphuric acid. It is a crystalline solid, which melts at 17.4°. Heated, it gives PH_3 and phosphoric acid, H_3PO_4. Hypophosphorous acid and its salts are powerful reducing agents.

Phosphorous Acid, H_3PO_3, is formed by the action of water on P_2O_3, or on PCl_3 (cf. p. 49), and also by the slow oxidation of phosphorus in moist air. A crystalline mass melting at 74°; deliquescent. Its solution absorbs oxygen from the air with formation of H_3PO_4, and is a good reducing agent. Decomposed by heat into PH_3 and H_3PO_4.

PHOSPHORIC ACID. $(HO)_3PO$.

Occurs abundantly in its salts (cf. phosphorus).

Formed by oxidation of phosphorus in the presence of water; by heating H_3PO_2 or H_3PO_3; by action of water on PCl_5, $POCl_3$ or P_2O_5.

Prepared: a. From bone ash by action of H_2SO_4. b. More easily pure by oxidation of phosphorus with HNO_3.

Properties. Transparent, crystalline, deliquescent mass. It is a strong tribasic acid, yielding three classes of salts—*phosphates.* Often called *orthophosphoric acid.* Heated to 200–300° it loses the elements of water and is converted into *pyrophosphoric acid*:—

$$2H_3PO_4 = H_2O + H_4P_2O_7.$$

Pyrophosphoric acid in solution changes when heated into the ortho-acid.

On heating the ortho or pyro-acid to redness, *metaphosphoric acid* is formed:—

$$H_4P_2O_7 = H_2O + 2HPO_3.$$

Metaphosphoric acid is also formed by dissolving P_2O_5 in cold water.

Metaphosphoric acid in solution changes slowly into the ortho-acid.

Glacial phosphoric acid is the meta-acid.

ARSENIC.

$As^{\text{III and V.}} = 75.$ $As_4 = 300.$

Occurs *free*—native arsenic. More often *combined;* with sulphur in realgar and orpiment; with metals, forming *arsenides ;* with sulphur and metals; as trioxide, and in arsenates.

Prepared: a. From its ores by sublimation.

b. From its oxide by reduction with charcoal:—

$$As_2O_3 + 3C = 3CO + 2As.$$

Properties. *Physical.*—Two allotropic modifications of arsenic are known. *Amorphous arsenic* is a black, lustreless mass or grey powder. Specific gravity, 4.71. Changes at 360° into the crystalline variety. *Crystalline* arsenic is a steel grey, brittle mass with metallic lustre. Specific gravity, 5.73. Both modifications sublime at 450° without previous fusion, yielding a lemon-yellow vapor of garlic-like odor. Arsenic is insoluble in all liquids except such as act upon it chemically.

Chemical.—Heated in air or oxygen, arsenic burns to the trioxide. Inflames in chlorine, burning to the trichloride. It unites with most elements directly. Unattacked by HCl; oxidized by HNO_3 or aqua regia to arsenic acid. Arsenic and most of its soluble compounds are exceedingly poisonous.

Uses. Metallic arsenic is chiefly used for the purpose of hardening lead in the manufacture of shot; in less amount in pyrotechny, and as fly poison.

History. Realgar and orpiment were known to the ancients. They were called *ἀρσενικόν.* Metallic arsenic was known to the alchemists.

HYDROGEN ARSENIDE. $AsH_3 = 78$.

Prepared: a. Pure by the action of H_2SO_4 or HCl upon an alloy of zinc and arsenic:—

$$As_2Zn_3 + 6HCl = 3ZnCl_2 + 2AsH_3.$$

b. Mixed with hydrogen, by the action of nascent hydrogen upon many arsenic compounds:—

$$As_2O_3 + 12H = 3H_2O + 2AsH_3.$$

Properties. Hydrogen arsenide or arsine is a colorless gas of peculiar and very disagreeable odor. Slightly soluble in water. It is extremely poisonous. Ignited it burns with a pale bluish flame forming H_2O and clouds of arsenic trioxide. Cold porcelain held in the flame receives a brown or black lustrous deposit of arsenic ; on heating the tube through which the gas is passing, arsine is decomposed and a *"mirror"* of metallic arsenic formed. (*Marsh's test.*)

ARSENIC AND THE HALOGENS.

$AsCl_3$, $AsBr_3$, AsI_3, AsI_2 and AsF_3 are known. They are analogous to the corresponding phosphorus compounds. The metallic character of arsenic is apparent in the fact that arsenic chloride, like other metallic chlorides, can be formed by the action of HCl upon the oxide.

OXIDES AND ACIDS OF ARSENIC.

ARSENIC TRIOXIDE. As_2O_3.

Occurs native in small amount as arsenolite.

Formed by burning arsenic or any compound of arsenic, and hence prepared in large quantities by roasting arsenical ores.

Properties. The trioxide condensed from vapor under ordinary conditions is a heavy white crystalline powder, of specific gravity 3.7, which sublimes at about 200° without fusion. When heated under pressure it melts and forms on cooling a transparent, vitreous, amorphous mass (arsenic-glass) of specific gravity 3.74, which after a time becomes opaque and porcelain-like, changing back to the crystalline modification.

Arsenic trioxide is slightly soluble in water (arsenious acid), much more soluble in HCl, and can be obtained from this solution in large crystals. Nascent hydrogen reduces it to AsH_3 (cf. p. 53); heated with charcoal it yields metallic arsenic. It is an energetic poison.

Used for the preparation of arsenic compounds: arsenic acid, which is largely employed in the manufacture of aniline colors; arsenical pigments (Schweinfurt-green, &c.), and other compounds of less importance; further, in glass-making; for preserving stuffed skins; as poison, &c. It is known commercially as white arsenic or arsenious acid.

Ortho-arsenic acid, H_3AsO_4, which is the complete analogue of phosphoric acid, is prepared by the oxidation of As_2O_3 by means of concentrated nitric acid.

Heated to 180° *pyro-arsenic acid* is formed:

$$2H_3AsO_4 = H_2O + H_4As_2O_7.$$

At 200° it is converted into *metarsenic acid*:—

$$H_4As_2O_7 = H_2O + 2HAsO_3.$$

Heated nearly to redness, the *pentoxide* is produced:

$$H_4As_2O_7 = 2H_2O + As_2O_5.$$

Treated with water As_2O_5, $HAsO_3$ and $H_4As_2O_7$ change into ortho-arsenic acid, H_3AsO_4.

Arsenious Acid, $(HO)_3As$, is probably contained in the solution of As_2O_3 in water. Its salts (arsenites) are stable compounds. Fowler's solution is potassium arsenite, and Scheele's green, copper arsenite. The alkaline arsenites are soluble in water, strong reducing agents and very poisonous.

ARSENIC SULPHIDES.

Arsenic Disulphide, As_2S_2, occurs naturally as *realgar* and may be prepared by fusing sulphur and arsenic in the proper proportions. An orange-red powder formerly much used as a pigment; is an ingredient of Indian or white fire.

Arsenic Trisulphide, As_2S_3, occurs naturally as *orpiment.* May be prepared by melting together sulphur and arsenic; and by precipitating a solution of As_2O_3 in hydrochloric acid, of an arsenite, or arsenic acid (in this case the As_2S_3 is mixed with sulphur) with H_2S. Yellow solid, soluble in alkaline sulphides with the formation of alkaline *sulpho-arsenites* (*e. g.* K_3AsS_3) ; in alkaline hydroxides and carbonates with the formation of an arsenite and a sulpho-arsenite.

Arsenic trisulphide is used as pigment (king's yellow); in dyeing with indigo; and mixed with lime and water as a depilatory in tanning.

Arsenic Pentasulphide, As_2S_5, is obtained by melting together the arsenic and sulphur in the proper proportions or by adding HCl to a solution of sodium sulpho-arsenate.

The *sulpho-arsenates* are formed by dissolving

As_2S_5 in an alkaline sulphide, or by treating As_2S_3 with an alkaline polysulphide.

ANTIMONY.

$Sb=120$.

Occurs occasionally native; chiefly, however, *combined* with sulphur (stibnite, Sb_2S_3); also with sulphur and metals in many minerals which almost always contain arsenic; more rarely as oxide.

Prepared from the natural sulphide: a. By heating with iron:—

$$Sb_2S_3+3Fe=3FeS+2Sb.$$

b. By roasting the ore to convert it into antimony oxide, and then reducing the oxide by heating it with charcoal:—

1. $Sb_2S_3+9O=3SO_2+Sb_2O_3$.
2. $Sb_2O_3+3C=3CO+2Sb$.

Properties. *Physical.* — Silver-white lustrous metal of coarsely crystalline structure; specific gravity, 6.71. It is hard and very brittle. Melts at 432° and distils at a white heat. *Chemical.*—Does not tarnish in air. On heating, it burns to Sb_2O_3; combines, like phosphorus, directly with the halogens. Very difficultly soluble in HCl; oxidized without being dissolved by HNO_3; dissolved by aqua regia.

Uses. Antimony is used in making tartar emetic and other preparations and is a constituent of several alloys—*e. g.* type metal, britannia metal, &c.

History. Its preparation was first described by Basil Valentine in the 15th century.

ANTIMONY AND HYDROGEN.

Hydrogen Antimonide, $SbH_3=123$, is formed in the same way as arsine; but never yet unmixed with hydrogen. It is a colorless gas of peculiar odor; burns with a greenish-white flame, forming H_2O and white clouds of Sb_2O_3. The flame gives spots on porcelain and a "*mirror*" is formed on heating the gas as in the case of arsine. The mirror is however readily distinguished from that of arsenic by its different appearance, its fusibility, less volatility and certain reactions.

ANTIMONY CHLORIDES.

Antimony Trichloride, $SbCl_3$, is obtained by heating antimony in a current of chlorine; in solution by dissolving antimony or its sulphide in HCl. It is a soft, deliquescent, crystalline mass (butter of antimony) melting at 73° and boiling at 223°. Decomposed by an excess of water with precipitation of SbOCl:—

$$SbCl_3+H_2O=SbOCl+2HCl.$$

Antimony Pentachloride, $SbCl_5$, is formed when $SbCl_3$ is saturated with chlorine. It is a colorless, volatile, fuming liquid. Decomposes easily into $SbCl_3$ and chlorine, and hence acts on many bodies energetically and in the same way as free chlorine.

$SbBr_3$, SbI_3 and SbF_3 are also known.

ANTIMONY OXIDES AND ACIDS.

Antimony forms three oxides: Sb_2O_3, Sb_2O_4 and Sb_2O_5. They are all acid-forming oxides; the first showing also feeble basic properties.

Antimony Trioxide, Sb_2O_3, occurs naturally in small quantities. It is formed by burning antimony in the air or oxidizing it with dilute nitric acid; also by precipitation of a boiling solution of $SbCl_3$ with sodium carbonate. With strong acids it gives unstable salts. With strong bases it also forms salts which are regarded as derivatives of $HSbO_2$. Heated, it gives antimony tetroxide, Sb_2O_4.

Salts of Antimony. There are two classes of salts: the [illegible] salts, derived from $Sb(OH)_3$; and the [illegible] salts, from $SbO.OH$.

Acids of Antimony. Acids similar to the phosphorus and arsenic acids are known: Antimonic acid, H_3SbO_4; pyroantimonic acid, $H_4Sb_2O_7$; and metantimonic acid, $HSbO_3$. Moderately heated, they are all converted into Sb_2O_5; at a higher temperature into Sb_2O_4.

ANTIMONY SULPHIDES.

Antimony Trisulphide, Sb_2S_3, occurs as *stibnite*, which is the most important ore of antimony. It is formed by melting together antimony and sulphur and is precipitated from antimony solutions by H_2S as an orange powder, which after fusion solidifies to a crystalline mass resembling stibnite. The natural sulphide is used as the source of antimony and its compounds, in pyrotechny and in making matches. The artificial compound is largely used in vulcanizing [illegible]ber, coloring it reddish-brown.

[illegible] sulphide dissolves in alkaline sulphides with [illegible]mation of sulpho-antimonites (cf. arsenic sulphides); and in alkaline polysulphides forming sulpho-antimonates.

Antimony Pentasulphide, Sb_2S_5, precipitated from a solution of a sulpho-antimonate by HCl, is a yellowish red powder. When heated decomposes into Sb_2S_3 and sulphur.

BISMUTH.

$Bi^{\text{III and V}}=207$.

Occurs chiefly *native;* more rarely *combined* with oxygen or sulphur.

Extracted from its ores by roasting and then smelting with the addition of iron, charcoal and slag. Bismuth is purified for pharmaceutical purposes by being melted with saltpeter.

Properties. *Physical.*—Bismuth is a hard, brittle metal, closely resembling antimony, from which it is distinguished by having a reddish tinge. Specific gravity, 9.8. Melts at 268° and expands as it solidifies.

Chemical.—Is unaltered in air at ordinary temperatures. On heating, it burns, forming a yellow oxide, Bi_2O_3. Unites directly with chlorine. Difficultly soluble in HCl; dissolves easily in HNO_3 or in aqua regia.

Bismuth forms no compound with hydrogen.

Used in certain fusible alloys which are employed for stereotyping, as solder, as safety plugs for boilers, &c. It is also used in making many pharmaceutical preparations and cosmetics, especially for *bismuth subnitrate.*

History. First described by Basil Valentine in the 15th century. Was for a long time confounded with antimony and zinc. Pott, 1739, first described its special properties.

COMPOUNDS OF BISMUTH.

Bismuth Chloride, $BiCl_3$, is formed by heating bismuth in a current of chlorine; in solution by dissolving the oxide in HCl or bismuth in aqua regia. It is a white, deliquescent mass, easily melted and distilled; is decomposed by an excess of water with the formation of *bismuth oxychloride,* BiOCl.

Bismuth forms similar compounds with bromine, iodine and fluorine.

Bismuth Oxides. The principal compound of bismuth and oxygen is *the trioxide,* Bi_2O_3, which occurs naturally as bismuth ochre; is prepared by burning bismuth or heating the nitrate or carbonate. It is a yellow powder. Used in making certain kinds of glass. The corresponding hydroxide, BiO.OH, is precipitated from bismuth solutions by ammonium or sodium hydroxide. On heating, it is converted into Bi_2O_3.

The other oxides are: Bi_2O_2 and Bi_2O_5.

Like antimony, bismuth forms two classes of salts: the bismuth salts and the bismuthyl salts.

Bismuth Nitrate, $Bi(NO_3)_3$, is obtained by dissolving bismuth in nitric acid. It is decomposed by water into white insoluble basic nitrates such as $Bi(OH)_2NO_3$ and $Bi(OH)(NO_3)_2$—the *subnitrate* of the pharmacopœia.

Bismuth Sulphide, Bi_2S_3, forms the mineral bismuthite and is precipitated from bismuth solutions by H_2S as blackish brown powder. Unlike the sulphides of arsenic and antimony, it is insoluble in alkaline sulphides.

BORON.

$B^{III}=11.$

Occurs only in combination in boric acid and borates.

Preparation. a. Boron is obtained as an amorphous powder by heating a mixture of boron oxide and sodium under a layer of common salt (to prevent access of air) :—

$$2B_2O_3+6Na=2Na_3BO_3+2B.$$

The molten mass is poured into very dilute hydrochloric acid in which the boron alone is insoluble.

b. By fusing amorphous boron or boron oxide with aluminium, and treating the mass when cold with hydrochloric acid, boron is obtained in crystalline form.

Properties. a. *Amorphous boron* is a brown infusible powder, which burns when heated in the air to B_2O_3. HNO_3 oxidizes it to boric acid. Brought into melted NaOH it gives sodium borate. It combines readily with chlorine, sulphur, nitrogen and other elements.

b. *Crystalline or adamantine boron* always contains aluminium. The crystals are brown, yellow or colorless, according to their purity, and have a hardness and lustre exceeded only by that of the diamond.

Specific gravity, 2.68. They are infusible, insoluble in acids and oxidize only on the surface when heated. They burn to BCl_3 when heated in chlorine.

COMPOUNDS OF BORON.

Hydrogen Boride, BH_3, is obtained (mixed with hydrogen) by the action of HCl on an alloy of boron and magnesium. Has not been prepared pure. It is a colorless gas of nauseating odor. Ignited, it burns with a fine green flame. Is decomposed by heat, with separation of boron.

Boron Trichloride, BCl_3, formed by direct union of boron and chlorine, is a colorless, fuming liquid, boiling at 18°. Decomposes with water into boric and hydrochloric acids.

Boron Trifluoride, BF_3, is a fuming gas made by treating a mixture of B_2O_3 and fluor-spar with H_2SO_4. It is very corrosive and carbonizes many organic substances. Exceedingly soluble in water, with which it forms boric acid and *fluoboric acid,* HBF_4.

Boron Trioxide, B_2O_3, is formed by burning boron or heating boric acid to redness. It is a brittle, glassy solid of specific gravity 1.8. Changes in moist air or in water to boric acid. It is non-volatile and hence decomposes at a red heat all salts whose acids are volatile, with the formation of borates. Most metallic oxides dissolve in it at a red heat, often imparting a characteristic color to the mass.

Boric Acid. *Orthoboric acid* $(HO)_3B$, occurs free in many volcanic regions, as in Tuscany, where it is contained in the steam which issues from volcanic vents. The steam is condensed in natural or artificial lagoons, and the acid crystallized from the water by evaporation. It is also prepared from *natural borax*, $Na_2B_4O_7$, by the action of HCl. Colorless, lustrous scales, soluble in water (1 : 26 at 19°; 1 : 3 at 100°) and slightly volatile with steam. Readily soluble in alcohol and this solution burns with a characteristic green-edged flame. It is a weak acid, forming unstable salts. Heated to 100° it loses water and is converted into *metaboric acid*, (HO)BO; at 140° a brittle, glass-like mass of *pyroboric acid*, $(HO)_2B_4O_5$, is obtained; at a red heat, B_2O_3.

The most important salt is $Na_2B_4O_7$, borax or sodium pyroborate.

Uses. Boric acid is used chiefly for the preparation of borax and also as a substitute for it. It is further employed as a flux in glass making; in making artificial gems; for impregnating candle wicks; in the preparation of preservatives for meat, &c.

CARBON.

$C^{IV}=12$.

Occurs *free* as diamond and graphite. Also more or less impure in coal. *Combined* in carbon dioxide and carbonates, and in all organic substances.

Properties. Three distinct modifications are known: the diamond, graphite and amorphous carbon. All three modifications of carbon are insoluble in all liquids, infusible and non-volatile, and burn in air or oxygen to carbon dioxide.

a. Diamond.—*Occurs* in certain alluvial deposits, usually loose in the sand. Ordinarily perfectly colorless and transparent, it is sometimes colored by slight impurities. It is the hardest substance known, and can be cut only by its own dust. Very brittle. Specific gravity, 3.5. It has a high lustre and refracts light powerfully. Heated between the poles of a strong electric battery (in an atmosphere free from oxygen) it swells up and is converted into a black, coke-like mass. Can be ignited in air or oxygen without difficulty. Is unattacked by any reagent. Lavoisier established its composition in 1776, by burning it in oxygen.

b. Graphite.—*Occurs* both foliated and massive. Is formed artificially by fusing iron with amorphous carbon; on cooling, part of the dissolved carbon crystallizes as graphite. It is a soft, steel-gray solid. Specific gravity, 2.3 (pure 2.09). In an atmosphere free from oxygen, it is unaltered at the highest temperatures. Burns in oxygen, but usually with more difficulty than the diamond. It is attacked by strong oxidizing agents (*e. g.* potassium chlorate and nitric acid).

Uses. Finely ground graphite mixed with clay is largely used for making "black-lead" pencils (originally solid slips of graphite) and crucibles. Graphite is also employed for stove-blacking; for polishing gun-powder; as a lubricant for machinery; and, on account of its conductivity for electricity, for coating the surface of electrotype moulds.

History. For many years graphite was supposed to contain lead and hence called black-lead or plumbago. Named from its use in pencils (γράφω).

c. Amorphous Carbon.—*Produced* by the imperfect combustion of organic matter, or by heating it out of contact with air. According to the source and properties several varieties are distinguished.

a. *Soot* or *lampblack* is prepared by burning substances rich in carbon (resins and oils) in a limited supply of air. It is nearly pure carbon, containing only small quantities of hydrocarbons. It is used in making India ink, black paint and printers' ink, and in calico printing.

b. *Gas carbon* is deposited in the upper portion of the retorts used in making gas as a result of the decomposition of gaseous hydrocarbons by heat. It is very hard and in its densest form has a specific gravity of 2.356. Nearly as pure as lamp-black. It is a good conductor of electricity.

c. *Charcoal,* produced by imperfect combustion (in heaps) or dry distillation of wood, contains about 85 per cent. of carbon and 3 per cent. ash (inorganic). The remainder is chiefly hygroscopic water, with small quantities of hydrogen and oxygen. Specific gravity, 0.145–0.17. It is very porous and absorbs gases and vapors freely (*e. g.* NH_3, H_2S), often taking up many times its own volume with evolution of heat, and also removes many substances from solution. It is used on this account as a disinfectant, and for purifying and decolorizing liquids. *Animal charcoal,* obtained by the carbonization of bones, possesses this property in a much higher degree than wood charcoal and is used largely in refining sugar.

Wood charcoal is an important fuel and reducing agent in metallurgical processes. It is also used in making gun-powder.

d. *Coke* is the residue left when the volatile matters of coal are driven off by heat. It is obtained by heating coal in special coke ovens and as a bye-product in making gas. It contains from 85–92 per cent. of carbon. It is dense and compact; and hence is ignited with difficulty and requires a blast or strong draft to keep it burning, but gives a very high temperature.

e. *Coal* is the result of a slow natural carbonization of vegetable matter. There are many different kinds containing more or less of the hydrogen, oxygen and nitrogen of the original substance. Cannel and boghead coals contain the largest proportion of volatile matters, and anthracite contains the least, while the bituminous coals lie between these extremes.

CARBON AND HYDROGEN.

Carbon forms a very large number of compounds with hydrogen—*hydrocarbons,* which, with their derivatives, form the subject of so-called *Organic Chemistry* or Chemistry of the Carbon Compounds. Only three of the simplest hydrocarbons will be here considered.

METHANE. $CH_4=16$.

Formed by the decomposition of organic matter out of contact with air. Hence occurs in coal mines (*fire-damp*) and petroleum wells; is formed at the bottom of stagnant pools (*marsh-gas*) and is always present in considerable quantities in coal gas.

Prepared by heating a mixture of sodium acetate and sodium hydroxide:—

$$NaO.OC.CH_3+NaOH=CH_4+Na_2CO_3.$$

Properties. Methane is a colorless, odorless gas, which can be condensed by cold and pressure. It is easily inflammable and burns with a faintly luminous flame to H_2O and CO_2. With twice its volume of oxygen (or ten times its volume of air) it forms an explosive mixture (coal mine explosions).

Methane is the first member of an homologous series of hydrocarbons, all of which can be represented by the general formula C_nH_{2n+2}.

ETHYLENE. $C_2H_4=28$.

Formed in the destructive distillation of coal, &c., and is consequently in coal gas, of which it is the most important illuminating constituent.

Prepared by withdrawing the elements of water from alcohol by means of strong sulphuric acid:—

$$CH_3.CH_2OH-H_2O=CH_2:CH_2.$$

Properties. Ethylene is a colorless gas of slight ethereal odor. Burns with a bright flame. Forms an explosive mixture with oxygen. Unites directly with an equal volume of chlorine, forming an oily liquid (hence called *olefiant gas*). It is the first member of a second series of hydrocarbons, which have the general formula C_nH_{2n}.

ACETYLENE. $C_2H_2=26$.

Formed by direct union of the elements at the temperature of the electric arc; by the incomplete combustion of many hydrocarbons (*e. g.* when a Bunsen burner "burns down"), or their decomposition at a red heat.

Properties. It is a colorless gas of very unpleasant odor, and burns with a luminous, smoky flame. It is the first member of a third series of hydrocarbons having the general formula C_nH_{2n-2}.

CARBON AND THE HALOGENS.

Carbon and the halogens cannot be made to unite directly. By the action of chlorine or bromine on hydrocarbons, *substitution products* are obtained in which the halogen has replaced hydrogen. In only a few cases can all of the hydrogen be replaced. The substitution usually proceeds by successive steps; thus from methane are obtained: CH_3Cl, CH_2Cl_2, $CHCl_3$ (*chloroform*) and CCl_4 (*carbon tetrachloride*). CCl_4 is a colorless, ethereal liquid, boiling at 77°.

CARBON AND OXYGEN.

CARBON MONOXIDE. $CO=28$.

Formed by the action of red-hot carbon (a) on carbon dioxide, and hence when fuel burns with too little air for complete combustion, $CO_2+C=2CO$; (b) on steam, $H_2O+C=CO+H_2$ (water gas); (c) on many metallic oxides, *e. g.* $ZnO + C = CO + Zn$. Also by heating many organic substances alone or with sulphuric acid.

Prepared by heating potassium ferrocyanide or oxalic acid with concentrated sulphuric acid :—

$$C_2O_4H_2 = CO_2 + H_2O + CO.$$

The CO_2 formed is removed by absorption by an alkaline solution.

Properties. *Physical.*—Colorless, odorless gas, somewhat lighter than air. Has been condensed by great pressure and cold. Very slightly soluble in water. It is very poisonous. *Chemical.*—Easily inflammable, burning with a pale blue flame to CO_2. In the heat, it is a strong reducing agent. Unites directly in sunlight with chlorine forming $COCl_2$.

History. First obtained by Lassone, 1776, by heating zinc oxide with charcoal. Composition determined by Cruikshank, 1800.

CARBON DIOXIDE. $CO_2 = 44$.

Occurs in the air in small quantities and in nearly all natural waters. It often escapes from the ground in volcanic regions in large quantities. It also occurs in liquid form in some minerals.

Formed: a. By burning carbon and organic substances and in the processes of respiration, fermentation and decay. b. From carbonates by heat or the action of acids.

Prepared: a. On a large scale by burning charcoal or coke or by heating limestone (calcium carbonate) :—

$$CaCO_3 = CaO + CO_2.$$

In the laboratory by the action of dilute hydrochloric acid on marble or chalk (calcium carbonate):

$$CaCO_3 + 2HCl = CaCl_2 + CO_2.$$

Properties. *Physical.*—Colorless gas of slightly pungent odor and acid taste. About 1.5 times heavier than air. On account of its weight, it often collects in old wells, coal pits, &c. (*choke-damp*) and can be poured from one vessel to another. Soluble in its own volume of water. This proportion is constant at all pressures, and hence (cf. Boyle's law) the amount of the gas which dissolves under pressure is simply proportional to the pressure. When the pressure is removed, the gas escapes with effervescence ("soda water," champagne, &c.). Carbon dioxide can be condensed (by 36 at. at 0°) to a colorless liquid. Freezes by spontaneous evaporation to a snow-like mass.

Chemical.—Incombustible and a non-supporter of combustion. Carbon dioxide is reduced to carbon by heated potassium, sodium or burning magnesium; to CO by carbon and certain metals (cf. p. 68). It is readily absorbed by alkaline solutions.

Uses. Carbon dioxide is used in making white lead, sodium carbonate and bicarbonate; in preparing "soda water" and mineral waters, for extinguishing fire and for many minor purposes. Liquid carbon dioxide is now manufactured for use as a fire-extinguisher, for charging liquids with the gas, and for the pressure thus rendered readily available.

Carbonic acid.—$(HO)_2CO$. Water in which CO_2 is dissolved, reddens litmus paper and is believed to contain carbonic acid in very dilute solution. On heating, the acid is decomposed and all of the CO_2 is expelled. Carbonic acid is *bibasic* and forms very stable salts—*carbonates* (*e. g.* Na_2CO_3). The carbonates of the alkalies are soluble in water; all other normal carbonates insoluble. They are readily decomposed by acids with evolution of carbon dioxide.

History. First distinguished from air by Van Helmont in the 16th century and called by him *gas sylvestre.* Black, 1755, found that it was absorbed by alkalies and called it *fixed air.* Lavoisier proved its composition in 1775. Carbon dioxide is commonly called carbonic acid gas.

THE NATURE OF FLAME.

Flame accompanies combustion only when gases or vapors are burning. Non-volatile substances and such as yield no combustible gases when heated (*e. g.* thoroughly carbonized charcoal) burn without flame.

The luminosity of a flame is influenced by the nature of the substances which it contains, by its temperature and by the density of the burning gases.

The luminosity of an ordinary lamp or gas flame is chiefly due to the presence of incandescent solid particles of carbon, which are separated from the burning hydrocarbons by their incomplete combustion in the interior of the flame.

The luminosity of any given flame is greater the higher its temperature or the denser the gases which form it. Thus the oxyhydrogen flame which is ordinarily non-luminous becomes bright when the gases burn under a pressure of twenty atmospheres; and a luminous gas flame is rendered non-luminous when the gas is diluted with air, nitrogen, carbon dioxide, hydrogen, &c. (Bunsen burner). If the gaseous mixture is first allowed to pass through a red-hot tube, the flame becomes again luminous.

Structure of Flame. The blowpipe flame.

CARBON DISULPHIDE. $CS_2=76$.

Formed like the dioxide by direct union of its elements, and *prepared* by passing the vapor of sulphur over red-hot charcoal.

Properties. *Physical.*—Carbon disulphide is a colorless, strongly refracting liquid of slight ethereal odor (when pure). Specific gravity, 1.29; very volatile and boils at 46°. Insoluble in water. Is an excellent solvent for many substances which do not dissolve in water, *e. g.* iodine, phosphorus, oils, fats, &c. *Chemical.*—Easily inflammable, burning with a blue flame to SO_2 and CO_2. Its vapor forms an explosive mixture with air. Dissolves in alkaline sulphides with the formation of *sulpho-carbonates* (*e. g.* $(NaS)_2CS$).

Commercial carbon disulphide has an intensely disagreeable odor from impurities which it contains.

Uses. On account of its solvent powers it is largely used for extracting fat from bones and wool, oils from oil seeds and olives, sulphur from its ore; as solvent for caoutchouc, phosphorus, iodine, &c.

Cyanogen, C_2N_2. Carbon unites with nitrogen only in the presence of metals and at a high temperature. Thus when refuse animal substances, such as blood, horn, hair, &c., are heated with potassium carbonate and iron, *potassium ferrocyanide,* $K_4Fe(CN)_6$

is formed. This is decomposed by heat, yielding potassium cyanide, KCN. From this salt other cyanides are readily made. Cyanogen is *prepared* by heating mercuric cyanide, $Hg(CN)_2$:

$$Hg(CN)_2 = Hg + C_2N_2.$$

Properties.—Colorless, poisonous gas, soluble in water. Burns with a purple flame, forming CO_2 and N.

Its name is due to the fact that several of its compounds have a blue color.

Hydrocyanic Acid, HCN, is *prepared* by the action of dilute acids on metallic cyanides. It is a volatile liquid, boiling at 26.5°. Very soluble in water. It has a very characteristic odor, suggesting that of bitter almonds. It is *extremely poisonous.* In its salts, the *cyanides,* the monad group CN plays the same part that chlorine plays in the chlorides.

SILICON.

$Si^{IV} = 28.$

Occurrence. Next to oxygen, silicon is the most abundant element. It occurs only in combination in silica (SiO_2) and silicates, which are the principal constituents of all rocks except the limestones. About one-quarter of the earth's crust is silicon.

Prepared by heating silicon fluoride (SiF_4) or sodium silico-fluoride (Na_2SiF_6) with sodium:—

$$Na_2SiF_6 + 4Na = 6NaF + Si.$$

Properties. On treating the result of the above reaction with water, silicon remains undissolved as a brown *amorphous* powder, which burns readily when heated in air to SiO_2; is insoluble in nitric or sulphuric acid; but dissolves easily in hydrofluoric acid to SiF_4, and in a hot concentrated solution of potassium hydroxide to K_4SiO_4 (potassium silicate). *Crystalline* silicon can be obtained by fusing sodium silico-fluoride with aluminium, or with sodium and zinc. The silicon dissolves, as it forms, in the aluminium or zinc and separates on cooling in dark, hard, lustrous crystals, which do not burn when heated; are insoluble even in hydrofluoric acid; but dissolve in boiling concentrated caustic alkali. Both modifications of silicon unite directly with chlorine when heated, forming $SiCl_4$.

History. Silicon was first obtained pure by Berzelius, 1823.

COMPOUNDS OF SILICON.

Silicon Hydride, $SiH_4=32$, is prepared (mixed with hydrogen) by treating an alloy of magnesium and silicon with hydrochloric acid:—

$$SiMg_2+4HCl=2MgCl_2+SiH_4.$$

It may be obtained pure by the action of sodium on silicon-triethylate.

It is a colorless, combustible gas. Spontaneously inflammable when diluted with hydrogen or subjected to diminished pressure. Burns brightly with the formation of clouds of SiO_2. Decomposes at a red heat into silicon and hydrogen; reacts at ordinary temperature with chlorine. Insoluble in water.

Discovered by Wöhler and Buff in 1857.

SILICON AND THE HALOGENS.

Silicon Chloride—$SiCl_4$, **Bromide**—$SiBr_4$, and **Iodide**—SiI_4, are obtained by the direct union of the elements or by the action of the halogen (gas or vapor) on a strongly heated mixture of (precipitated) silica and charcoal. These compounds are decomposed by water into the halogen acid and H_4SiO_4.

Silicon Fluoride, SiF_4, is formed by the action of hydrofluoric acid on silicon, silica or silicates :—

$$SiO_2 + 4HF = SiF_4 + 2H_2O.$$

Prepared by warming a mixture of sand or powdered glass and fluor spar (CaF_2) with concentrated sulphuric acid :—

$$2CaF_2 + SiO_2 + 2H_2SO_4 = SiF_4 + 2CaSO_4 + 2H_2O.$$

Properties.—Silicon fluoride is a colorless gas of very sharp odor. It fumes in moist air and is decomposed instantly by water with the separation of gelatinous silicic acid.

Hydrofluosilicic Acid, or *silico-fluoric acid*, H_2SiF_6, is formed when silicon fluoride is decomposed by water :—

$$3SiF_4 + 4H_2O = 2H_2SiF_6 + (HO)_4Si.$$

It is known only in aqueous solution, being decomposed when this is evaporated into SiF_4 and HF. The saturated solution is a colorless, fuming and very acid liquid. The salts—*silico-fluorides*—are mostly soluble in water, notable exceptions being the salts $BaSiF_6$, K_2SiF_6 and Na_2SiF_6.

SILICON DIOXIDE. SiO_2.

Occurs abundantly as quartz (rock-crystal, amethyst, chalcedony, jasper, &c.), opal, sand and kiesel-guhr. It is also contained in the stems of cereals and is found dissolved in many hot springs.

Formed by burning silicon or silicon hydride, or by heating silicic acid.

Properties. Natural, crystallized quartz is colorless or variously colored by slight impurities and is very hard. Specific gravity, 2.6. It melts only at the temperature of the oxyhydrogen blowpipe; is insoluble in water and all acids except HF; is unattacked by boiling solutions of caustic alkalis, but forms a glassy mass of alkaline silicate (water glass) when fused with alkaline hydroxides or carbonates.

Artificially prepared silicon dioxide is a white, voluminous powder. Both this and the natural amorphous varieties differ from the crystalline by being soluble in solutions of the alkaline carbonates and hydroxides.

Uses. Some of the natural varieties are used as gems and ornaments; rock crystal for optical lenses; kiesel-guhr in making dynamite; sandstone for building purposes; sand in making glass, porcelain, and mortar.

SILICIC ACID.

Ortho-silicic Acid, $(HO)_4Si$, is probably contained in the precipitate formed when SiF_4 is decomposed by water. Obtained in solution by pouring a solution of an alkaline silicate (*e. g.* $(NaO)_4Si$) into dilute hydrochloric acid, and separated from the excess of HCl and the NaCl formed in the reaction by *dialysis*. This solution may be concentrated by boiling till it has a strength of 14 per cent.; then coagulates to a jelly, which on drying has the composition $(HO)_2SiO$, *meta-silicic acid*. The gelatinous acid is obtained at once when hydrochloric acid is poured into the alkaline silicate solution. Ortho-silicic acid is a weak, unstable acid and few of its salts are known. Most of the natural silicates are derivatives of meta-silicic acid or more or less complex hypothetical *poly-silicic acids*, which are derived from the ortho and meta-acids by elimination of the elements of water from two or more molecules (cf. phosphoric and arsenic acids). Only silicates of the alkalis are soluble in water; most of the others are insoluble in acids; but some are dissolved by hydrochloric acid with the separation of gelatinous silicic acid. All silicates are decomposed by hydrofluoric acid or by fusion with alkaline hydroxides or carbonates.

APPENDIX.

LIST OF ELEMENTS, SYMBOLS AND ATOMIC WEIGHTS.

Element.	Symbol.	Atomic Weight.
Aluminium,	Al	27.04
Antimony, (Stibium)	Sb	119.6
Arsenic,	As	74.9
Barium,	Ba	136.86
Beryllium,	Be	9.08
Bismuth,	Bi	207.5
Boron,	B	10.9
Bromine,	Br	79.76
Cadmium,	Cd	111.7
Cæsium,	Cs	132.7
Calcium,	Ca	39.91
Carbon,	C	11.97
Cerium,	Ce	141.2
Chlorine,	Cl	35.37
Chromium,	Cr	52.45
Cobalt,	Co	58.74
Columbium,	Cb	93.7
Copper, (Cuprum)	Cu	63.18
Didymium,	Di	142.1
Erbium,	E	166
Fluorine,	F	19.06
Gallium,	Ga	69.9
Germanium,	Ge	72.31
Gold, (Aurum)	Au	196.2
Hydrogen,	H	1
Indium,	In	113.4
Iodine,	I	126.54
Iridium,	Ir	192.5
Iron, (Ferrum)	Fe	55.88
Lanthanum,	La	138.5
Lead, (Plumbum)	Pb	206.39
Lithium,	Li	7.01
Magnesium,	Mg	23.94
Manganese,	Mn	54.8
Mercury, (Hydrargyrum)	Hg	199.8
Molybdenum,	Mo	95.9
Nickel,	Ni	58.56
Nitrogen,	N	14.01
Osmium,	Os	195
Oxygen,	O	15.96
Palladium,	Pd	106.2
Phosphorus,	P	30.96
Platinum,	Pt	194.34
Potassium, (Kalium)	K	39.03
Rhodium,	Rh	104.1
Rubidium,	Rb	85.2
Ruthenium,	Ru	103.5
Scandium,	Sc	43.97
Selenium,	Se	78.87
Silicon,	Si	28
Silver, (Argentum)	Ag	107.66
Sodium, (Natrium)	Na	23
Strontium,	Sr	87.3
Sulphur,	S	31.98
Tantalum,	Ta	182
Tellurium,	Te	125
Thallium,	Tl	203.7
Thorium,	Th	231.96
Tin, (Stannum)	Sn	117.35
Titanium,	Ti	50.25
Tungsten, (Wolfram)	W	183.6
Uranium,	U	239.8
Vanadium,	V	51.1
Ytterbium,	Yt	172.6
Yttrium,	Y	89.6
Zinc,	Zn	65.1
Zirconium,	Zr	90.4

Natural System of the Elements.—Mendelejeff's Table.

SERIES.	GROUP I. — R_2O.	GROUP II. — RO.	GROUP III. — R_2O_3.	GROUP IV. RH_4, RO_2.	GROUP V. RH_3, R_2O_5.	GROUP VI. RH_2, RO_3.	GROUP VII. RH, R_2O_7.	GROUP VIII. — RO_4.
1	H = 1							
2	Li = 7	Be = 9.1	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27	Si = 28	P = 31	S = 32	Cl = 35.4	Fe = 56, Ni = 58.6, Co = 58.8, Cu = 63.
4	K = 39	Ca = 40	Sc = 44	Ti = 50	V = 51	Cr = 52.4	Mn = 55	
5	(Cu) = 63	Zn = 65	Ga = 70	Ge = 72	As = 75	Se = 79	Br = 80	Ru = 103.5, Rh = 104, Pd = 106, Ag = 108.
6	Rb = 85	Sr = 87	Y = 89	Zr = 90	Cb = 94	Mo = 96	— = 100	
7	(Ag) = 108	Cd = 112	In = 113	Sn = 117	Sb = 120	Te = 125	I = 126.5	
8	Cs = 133	Ba = 137	La = 138.5	Ce = 141	Di = 142	—	—	— — —
9	(—)	—	(Er. 166)	—	—	—	—	Ir = 193, Pt = 194, Os = 195, Au = 196.
10	—	—	Yb = 173	—	Ta = 183	W = 184	—	
11	(Au) = 196	Hg = 200	Tl = 204	Pb = 206.4	Bi = 207.5	—	—	
12	—	—	—	Th = 232	—	U = 240	—	— — —

THE METRIC SYSTEM.

MEASURES OF LENGTH.

The unit is the **Meter** (m)=39.37 inches.

Subdivisions.		*Multiples.*	
Millimeter (mm)=0.001	meter.	Decameter (Dm)=10	meters.
Centimeter (cm)=0.01	"	Hectometer (Hm)=100	"
Decimeter (dm) =0.1	"	Kilometer (Km)=1000	"

MEASURES OF AREA.

Square meter (qm) with subdivisions and multiples as above. The **Are** (a)=100 qm; the **hectare**=100 a=2.4714 acres.

MEASURES OF VOLUME.

The **Liter** (l)=1 cubic decimeter (cdm)=1.0567 U. S. quarts. The decimal subdivisions and multiples have the same prefixes as the measures of length.

WEIGHTS.

The unit is the **Gram** (g), which is the weight of 1 cubic centimeter (1ml) of pure water at 4° C. The gram=15.432 grains or 0.03527 oz. av. The decimal subdivisions and multiples of the gram are designated by the same prefixes as those used for the measures of length.

DATA FOR CONVERTING ENGLISH MEASURES AND WEIGHTS INTO THOSE OF THE METRIC SYSTEM.

1 inch=2.54 cm.
1 foot=3.058 dm.
1 yard=0.9143 m.
[illegible]9 Km.
[illegible]
[illegible]=0.947 l.
[illegible]=3.785 l.

1 grain=64.8 mg.
1 oz. Troy=31.1035 g.
1 pound Troy=373.23 g.
1 oz. avoir.=28.35 g.
1 pound avoir.=453.59 g.

THERMOMETER SCALES.

The scales are: *Centigrade* (C.), for most scientific work; *Fahrenheit* (F.), for ordinary purposes in England and America; *Réaumer* (R.), in Europe.

On each scale the two "fixed points" are determined at which the mercury stops when the thermometer is placed in melting ice (freezing point), and in steam of water boiling under normal pressure (boiling point). These points are marked:—

Freezing point,	0°C.	0°R.	32°F.
Boiling point,	100°	80°	212°

Between the fixed points the scale is divided into equal degrees, the number of which in each case is indicated by the figures above. Thus 100°C.=80°R.=180°F., or in the simplest terms, 5°C.=4°R.=9°F.

Hence the following formulas serve for the conversion of the readings of any one scale into those of the others:—

$$t^\circ\mathrm{C.}=\tfrac{4}{5}t^\circ\mathrm{R.}=(\tfrac{9}{5}t+32)^\circ\mathrm{F.}$$
$$t^\circ\mathrm{R.}=\tfrac{5}{4}t^\circ\mathrm{C.}=(\tfrac{9}{4}t+32)^\circ\mathrm{F.}$$
$$t^\circ\mathrm{F.}=\tfrac{5}{9}(t-32)^\circ\mathrm{C.}=\tfrac{4}{9}(t-32)^\circ\mathrm{R.}$$

Comparison of the three scales.

C.	*F.*	*R.*	*C.*	*F.*	*R.*
−40	−40	−32	60	140	48
30	22	24	70	158	56
20	4	16	80	176	64
10	+14	8	90	194	72
0	32	0	100	212	80
+10	50	+8	200	392	160
20	68	16	300	572	240
30	86	24	400	752	320
40	104	32	500	932	400
50	122	40	1000	2120	800

CHEMICAL CALCULATIONS.

1. **Quantitative Composition.** Molecular formulas furnish the data for the calculation of the quantitative composition of substances; for the symbols which compose the formulas express not only kinds of matter, but parts by weight. *E. g.* $H_2SO_4=98$; $H_2=2$; $S=32$; $O_4=64$. Hence the proportions by weight are: hydrogen, $\frac{2}{98}$; sulphur, $\frac{32}{98}$; oxygen, $\frac{64}{98}$. From the proportions thus obtained, the percentage composition or the weight of a constituent in a given weight of substance can easily be found.

2. **Reactions.** The amounts of the substances which enter into and result from a reaction are readily calculated from the equation which represents the reaction. *E. g.* $Zn+H_2SO_4=ZnSO_4+H_2$ expresses the fact that 65 parts of zinc and 98 parts of sulphuric acid yield 161 parts of zinc sulphate and 2 parts of hydrogen. Hence to obtain the weight of zinc sulphate which will be formed from any given weight of zinc, the following proportion is employed:

$$65 : 161 :: \text{wt. of zinc} : \text{wt. of } ZnSO_4.$$

3. **Conversion of the weight of a gas into normal volume.** This is readily accomplished when the weight of one liter of the gas under standard conditions is known. The weight of a *normal liter* of any gas equals the weight of a normal liter of hydrogen (0.0896) multiplied by one-half the molecular weight of the gas (cf. p. 8).

4. **Measuring Gases.** The volume occupied by a given weight of a gas depends upon the pressure, the temperature and the presence or absence of moisture. The *normal volume* is that of dry gas measured at 0°C. under a pressure equal to that of 760 mm. of mercury. A gas measured under other conditions than these can be reduced to normal volume by the following formulas, in which V_0 is the normal volume and V the volume measured at the temperature $t°$ and under the pressure B (barometer reading):—

Since $V_0 : V :: B : 760$ (cf. Boyle's law), $V_0=V\frac{B}{760}$.

Also, since $V_\circ : V :: 273 : 273+t^\circ$ (cf. Charles' law), $V_\circ = V\frac{273}{273+t} = V\frac{1}{1+0.00366t}$.

If the gas is measured over water, the tension of the aqueous vapor at t° (w) causes the gas to occupy a greater volume and must be deducted from B. Hence the complete formula for finding *the normal volume of a gas measured over water at t° and under the pressure B*, is:

$$V_\circ = V\frac{273(B-W)}{760(273+t)} = V\frac{B-W}{760(1+0.00366t)}.$$

To find the volume which a given normal volume will occupy when collected over water at t° and under the pressure B:

$$V = V_\circ\frac{760(273+t)}{273(B-W)} = V_\circ\frac{760(1+0.00366t)}{B-W}.$$

To convert a known volume of a gas into its weight: Reduce, if necessary, to normal volume and multiply by the weight of one normal liter of the gas.

Tension of aqueous vapor at different temperatures expressed in millimeters of mercury.

At C°.	Tension.	At C°.	Tension.
0	4.6	21	18.495
5	6.534	22	19.659
10	9.165	23	20.888
15	12.699	24	22.184
16	13.536	25	23.550
17	14.421	30	31.548
18	15.357	35	41.827
19	16.346	40	54.906
20	17.391	50	91.982

Determination of the Specific Gravity (Air = 1) of Gases and Vapors from their Molecular Weights. Since the specific gravity (H = 1) is one-half the molecular weight, and hydrogen is 14.44 times lighter than air, the specific gravity (air = 1) is found by dividing the molecular weight by 28.88.

Conversely, the molecular weight of a gas or vapor may be obtained from its specific gravity (air=1) by multiplying this by 28.88.

TYPICAL PROBLEMS.

Determine :—

1. The percentage composition of zinc sulphate.

2. The weight of zinc sulphate which can be made from 5 grams of zinc.

3. The weight of zinc and of pure sulphuric acid necessary to make 5 grams of zinc sulphate.

4. The normal liters of hydrogen which can be obtained from sulphuric acid by the use of 5 grams of zinc.

5. The weight of zinc necessary to make 5 normal liters of hydrogen.

6. The volume of oxygen, measured over water at 20° and 750 mm pressure, which can be made from 5 grams of potassium chlorate.

7. The weight of potassium chlorate necessary to make 5 liters of oxygen, measured dry at 20° and 750 mm pressure.

8. The volume occupied by the oxygen of the last problem when measured over water at 15° and 765 mm pressure.

9. The specific gravity (air=1) of HCl.

10. The molecular weight of water; specific gravity of steam (air=1)=0.6235.

Specific Gravities of the Non-metallic Elements and some of their Compounds.

Solid. Water = 1.00.		Liquid. Water = 1.00.		Gases, vapors, cont'd.	
P (ord)	1.84	N_2O	0.937	HCN	0.948
P (red)	2.15	CO_2	0.947	CO	0.967
S (mcl)	1.96	H_2O	1.000	N	0.971
S (rh)	2.07	CS_2	1.29	C_2H_4	0.978
C (graphite)	2.33	Cl	1.33	NO	1.039
C (diamond)	3.52	SO_2	1.43	O	1.105
Si	2.49	HNO_3	1.53	H_2S	1.178
B	2.68	PCl_3	1.61	HCl	1.26
Se (amorph.)	4.48	H_2SO_4	1.84	N_2O	1.527
Se (cryst.)	4.8	Br	3.15	CO_2	1.524
As (amorph.)	4.71	**Gases and vapors.**		C_2N_2	1.806
As (cryst.)	5.73	**Air = 1.00.**		SO_2	2.24
I	4.95	H	0.06926	Cl	2.45
Te	6.18	CH_4	0.557	CS_2	2.645
Sb	6.7	NH_3	0.589	P	4.35
Bi	9.8	H_2O	0.623	Br.	5.54
NH_3	0.636	C_2H_2	0.92		

Temperatures and Pressures at which Certain Gases are Liquefied.

	Temp. ° C.	Pressure in Atmos.		Temp. ° C.	Pressure in Atmos.
SO_2	− 8	1	CO_2	− 80	1.1
	+17.8	2.8		+ 15.5	52
Cl	−33.6	1	N_2O	− 87.2	1
	+12.5	8.5	NO	− 11	104
NH_3	−38.5	1	O	−181.4	1
	+20.3	8	CO	−190	1
H_2S	−70	1.09	Air	−192	1
	+11.1	14.6	N	−193.1	1
			H	−140	652

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